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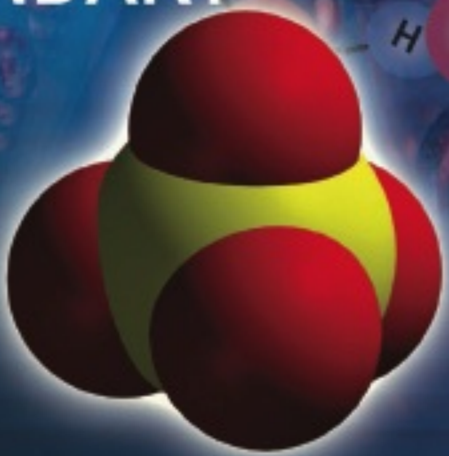
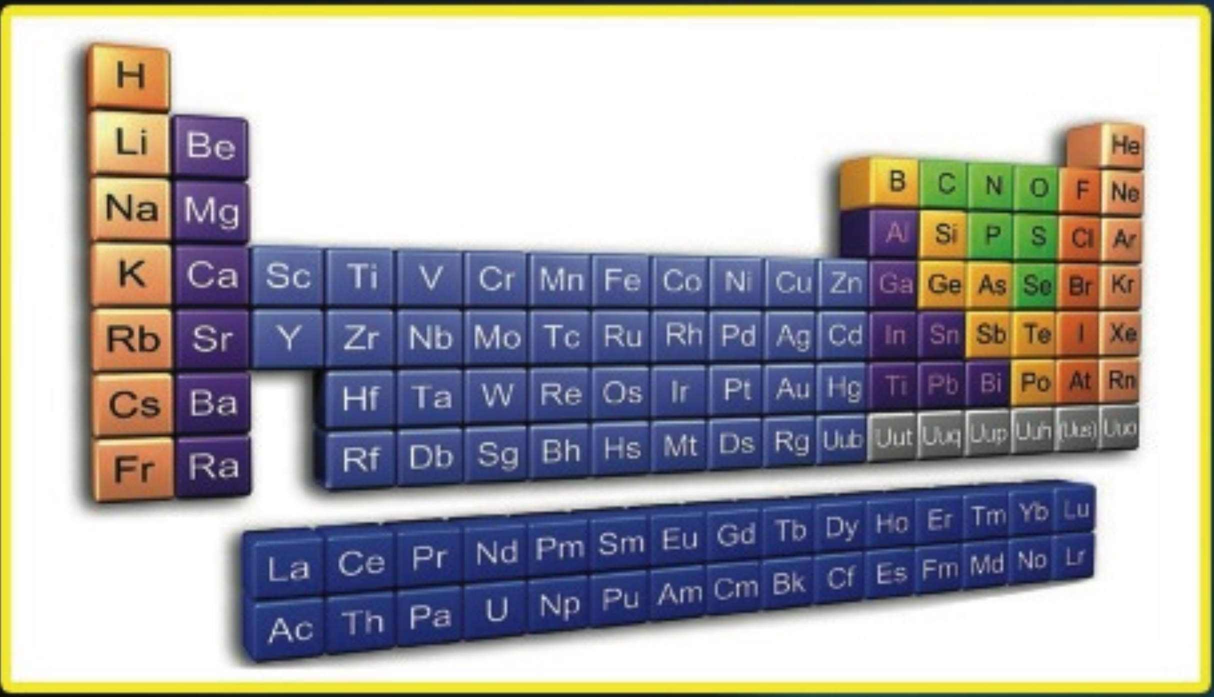
CHEMISTRY

SECOND SECONDARY



2016 - 2017

غير مصرح بتداول هذا الكتاب خارج وزارة التربية والتعليم



Arab Republic of Egypt
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CHEMISTRY

FOR Second Secondary

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2016-2017

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المقدمة

في إطار تطوير التعليم لمواكبة المتغيرات العالمية والمحلية واستكمالا للجهود الحثيثة التى تقوم بها وزارة التربية والتعليم للارتقاء بمستوى المناهج الدراسية وربطها بالمجتمع والبيئة فقد كلف الأستاذ الدكتور وزير التربية والتعليم نخبة من أساتذة الجامعات المتخصصين بالتعاون والتنسيق مع موجهي الكيمياء بالوزارة وبمشاركة مؤلفي الكتاب السابق لإعادة تقييم ومراجعة المحتوى العلمي لمادة الكيمياء للصف الثانى الثانوى.

ولقد قامت اللجنة المكلفة بإجراء التعديلات والإضافات اللازمة التى أدت إلى :

(1) التخلص من التكرار والحشو غير المبرر واستبعاد الأجزاء التي سبق للطالب دراستها وإعادة صياغة بعض أجزاء الكتاب بطريقة منطقية متسلسلة ومنظمة.

(2) إضافة بعض المفاهيم والتطبيقات لمواكبة الاتجاهات العلمية الحديثة.

(3) ربط موضوعات الدراسة بالحياة اليومية وتأثيراتها البيئية وتطبيقاتها الصناعية.

(4) الاهتمام بالمعالجات الرياضية في فهم بعض الموضوعات بهدف تقوية الجوانب الفكرية.

(5) إدخال بعض الموضوعات التى تتيح للطالب إجراء تجاربها معمليًا لاكتساب مهارات عملية.

(6) إعداد بعض الأشكال التوضيحية وتوظيفها لخدمة المفاهيم العلمية.

(7) تحديد الأهداف المرجوة من دراسة كل فصل من فصول الكتاب وضعت في مقدمته لتعطى مؤشراً للطالب والمعلم على مدى ما حققه.

(8) تنوع التقييم ليتضمن قياس المستويات المختلفة للتعليم.

والكتاب فى صورته الحالية يحتوى على أربعة أبواب تتكامل وتترابط فيما بينها وتعكس تناغمًا مع محتويات كتب الكيمياء في المناهج العالمية وتشتمل على تطبيقات صناعية وبيئية مفيدة وتتضمن اهتمامًا واضحًا بتنمية قدرات الفهم والابتكار وتتمشى مع المعايير القومية التي وضعتها الوزارة لتطوير منهج الكيمياء نتمنى أن يكون هذا الكتاب في صورته الجديدة مصدرًا مفيدًا للعلم والمعرفة في مجال الكيمياء وأن يحقق الغاية المرجوه وأن يكون خير معين لطلابنا الذين نتمنى لهم النجاح والتوفيق.

لجنة التطوير

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Chapter One

Atomic Structure

Objectives

At the end of this chapter, the student should be able to :

- Recognize the historical background of atomic structure.
- Describe the properties of cathode rays.
- Discuss the Rutherford s atomic model.
- Recognize the Bohr's atomic model.
- Define the reasons of the inadequacy of Bohr's model.
- Construes the modification introduced by the modem atomic theory.
- Explain the concept of electron cloud and orbitals.
- Define the four quantum numbers.
- Distribute electrons of any atom considering the building up principle and Hund's rule, and Pauli's exclusion principle .
- Appreciate the efforts of scientists in the development of chemistry

ATOMIC STRUCTURE

Introduction:

Long time ago man was asking about the nature of matter and its structure.

To answer this question, the Demokrats (Greek Philosopher) imagined the possibility of dividing any piece of matter to smaller parts , then dividing those parts into smaller particles and further into smaller ones and so on, until we reach an individable fragment. They named it an "atom" (In the Greek language "A" means "no" and "TOM" means "divide") .

- Aristotle, in the 4th century B.C. rejected the concept of "the atom", and believed that all matters, whatever their nature, are composed of four components, they are: water, air, dust, and fire. It was believed that cheap metal as iron or copper can be changed into precious ones as gold by changing the ratio of the constituents of the four components: This miserable idea Caused a retard of development in chemistry science for more than 1000 years.
- The Irish scientist, Boyle, in the year 1661, refused Aristotle's idea about the nature of substances and gave the first definition of the element as: " a pure simple substance that cannot be changed to simpler forms by the traditional chemical methods".
- At the beginning of the nineteenth century, the English scientist, John Dalton carried out many researches and experiments and stated the first theory about the atomic structure.

Dalton's atomic theory

Dalton announced in 1803 his atomic theory, in which he postulated the following:-

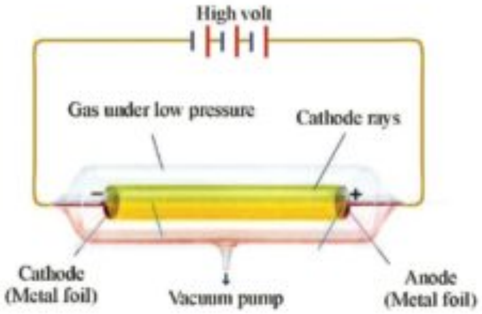
1. The element is composed of very minute particles which are named atoms.
2. Each element is composed of very minute individable solid atoms.
3. Atoms of the same element are similar in mass but differ from atoms of other elements.
4. Compounds are formed by the combination of atoms of different elements in a simple numerical ratio.

Discovery of cathode rays

In 1897, many experiments were carried out on the electric discharge through gases. It was known that gases do not conduct electricity under normal conditions of pressure and temperature.

However, they could conduct electricity in discharge tubes with a very low pressure. Then a potential difference of about 10,000 volts was applied through the space between the electrodes . A stream of invisible rays was emitted from the cathode causing a fluorescent glow on the tube wall. These rays were named the "cathode rays", and afterward it was known that they are composed of fine particles named by electrons.

Properties of cathode rays



1- They consist of very fine negatively charged particles.
2- They move in straight lines.
3- They have a thermal effect.
4- They are affected by both electric and magnetic fields.
5- They do not differ either in behavior or in nature if the material of the cathode or the used gas are changed. This is a strong evidence that it is a fundamental constituent of any matter.

The Thomson's model of the atom

In 1897, Thomson suggested a new atomic model of the atom. He considered the atom as "a sphere of a uniform positive electricity in which a number of negatively charged electrons are embedded in to make the atom electrically neutral".

Rutherford's atom

In 1911, two of Rutherford's students, Geiger and Marsden, performed a famous laboratory experiment using the illustrated apparatus fig (1-2).

The experiment.

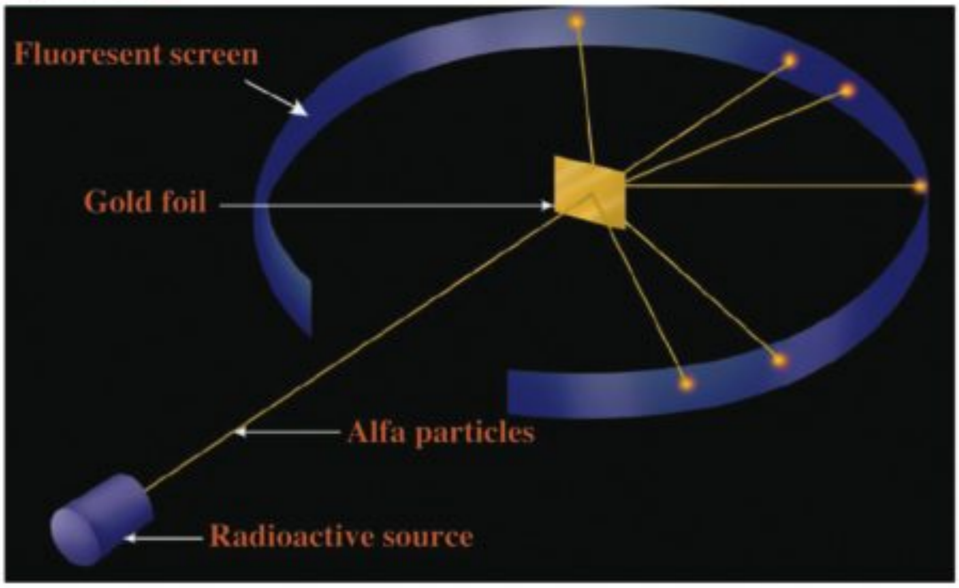


Fig: 1-2

1- Rutherford allows alpha particles to collid a metal sheet lined with a layer of zinc sulphide which glows at the site of collision with (α) - particles in absence of gold foil. It was possible to define the location and number of alpha particles (α) that hit the metal sheet from counting the glows.

2 - A very thin gold foil was placed between the beam of alpha particles and the metal sheet.

Rutherford recorded his observations and reached the following conclusions;

Observation	Conclusion
1- The majority of alpha(α) -particles hit the same places in which they appeared before placing the gold foil.	1- Most of the atomic volume is an empty space and the atom is not uniformly dense as proposed by Thomson and Dalton.
2- A very small percentage of α - particles did not penetrate the gold foil and reflected back where some flashes appeared in front of the foil.	2- The atom must contain a tiny part of a very high density (was named the nucleus).
3- Some of (α) -particles penetrated the foil but were deflected.	3- The dense part of the atom where most of the mass is present, appears to have a similar positive charge to that of α - particles.

Fig (1-3) A diagram explains Rutherford's experiment.

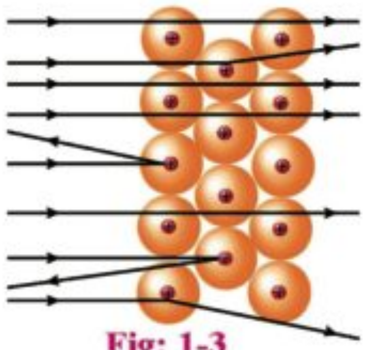


Fig: 1-3

Rutherford's atomic model

On the basis of his experiment and from the experiments of other scientists, Rutherford designed his atomic model as follows:-

1- The atom: Despite its extremely small size, it has a complicated structure that resembles (the sun) around which electrons (planets) orbit.

2. The nucleus: is much smaller than the atom; and there is a vast space between the nucleus and the orbits of electrons (i.e. the atom is not uniformly dense). Most of the atomic mass and the positive charge are concentrated in the nucleus.

3. The electrons :

a- These have negligible mass to that of the nucleus.

b- The sum of negative electric charges of electrons equals the nuclear positive charge (i.e. the atom is electrically neutral).

C- Electrons travel around the nucleus in special orbits at a tremendous speed despite the mutual attraction between them and the nucleus. The attraction force is overcome by another force that equals it in quantity and opposes it in direction namely the centrifugal force. But Rutherford's theory does not explain the system at which electrons revolve the nucleus.

Atomic spectra and its explanation(Bohr's atomic theory)

The study of atomic spectra is considered the key which solved the puzzle of atomic structure. That was the work of the Danish scientist Niels Bohr (1913) upon which he was rewarded the Nobel prize in 1922.

Atomic emission spectra

On heating atoms of a pure element in gaseous or vapour state to high temperature or exposing them to low pressure inside electric discharge tube, they emit radiation known as line spectrum. On examining this radiant light by a spectroscope, we observe a group of small number of restricted coloured lines separated by dark areas so it's called line spectrum. This phenomenon can't be explained by physics scientists. It was found experimentally that the spectral lines are essential characteristic for each element i.e. there are no two elements that have the same spectral lines.

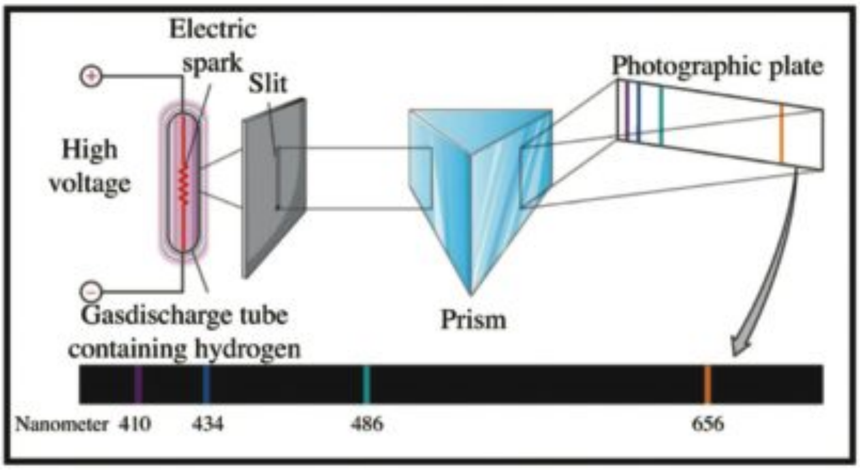


Fig: 1-4 Line spectrum for Hydrogen

Bohr's atomic model

Studying the line spectra of hydrogen atoms, Bohr was able to reach his atomic model

Bohr Postulates

Bohr adopted some of Rutherford's postulates about atomic structure;

1. A positively charged nucleus exists in the center of the atom.
2. The number of negative electrons equals the number of positive charges which the nucleus carries.
3. While the electron orbits the nucleus, a centrifugal force arises which is compensated by the attraction force of the nucleus for the electron.

Bohr added the following hypotheses to those of Rutherford's.

4. Electrons orbit, the nucleus in a rapid movement without emission or absorption of any amount of energy.

5. Electrons orbit the nucleus only in definite allowed energy levels.

They cannot be found at intermediate distances.

6- Each electron in the atom has a definite amount of energy depending on the distance between its energy level and the nucleus; the energy of any level increases as its radius increases.

Each energy level is expressed by a whole number called the principal quantum number.

7. The electron remains in the lowest allowed energy level in its ground state. However; if it acquires an amount of energy (termed quantum) by heating or by electric discharge, the electrons become excited and jump (temporarily) to some higher energy level depending on the absorbed quantum . The excited electron in the higher energy level is then unstable and soon returns to its original level losing the same quantum of energy which it absorbed during its excitation, but in the in form of radiant light of a definite wave length (or frequency) producing a characteristic spectral line.

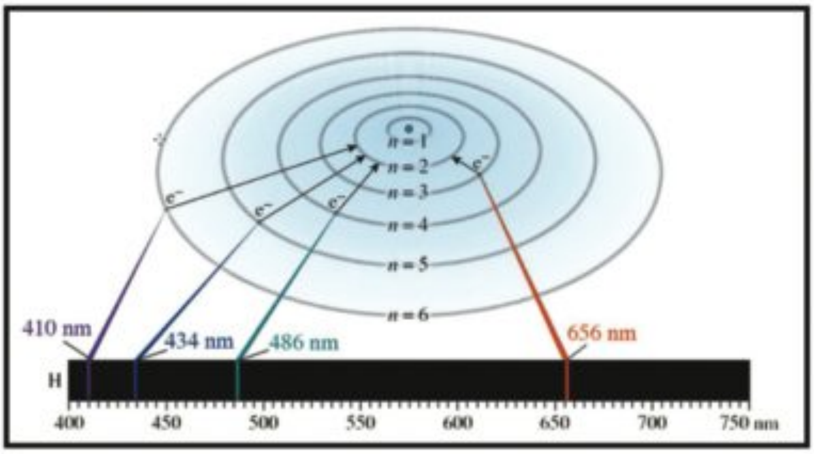


Fig: 1-5 The line spectrum for hydrogen is comosed of four colred line

8, The multitude of atoms absorb different amount of energy, then radiate their energies producing spectral lines. These lines correspond to the energy levels from which their electrons are transmitted back to the ground state (explanation of the spectral lines in the hydrogen atom).

The following notes must be taken into consideration

1. The quantum is the amount of energy absorbed or emitted when an electron jumps from one energy level to another.

2. Bohr's calculations of the energy levels radii and of the energy of each level revealed that the difference in energy between them is not equal, i.e. the energy gap decreases further from ' the nucleus. This means that the quantum of energy required to transfer an electron from one energy level to another is not equal.

3. The electron does not move from its energy level to another unless the energy absorbed or emitted is equal to the energy difference between the two levels i.e. one quantum. It is impossible for an electron to move from its orbit if the energy absorbed or emitted is less than one quantum (i.e. there is no half quantum for instance).

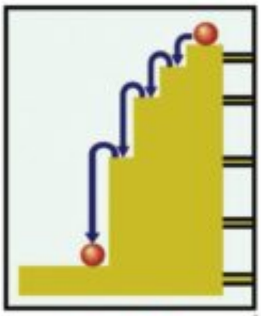


Fig: 1-6

Bohr's atomic model succeeded to a great extent in the following :

1. It explained the hydrogen atom spectrum.
2. It introduced the idea of quantized energy state for electrons in the atom.

Inadequacy of Bohr's atomic model

Despite the great efforts of Bohr to construct his atomic model, the quantitative calculations of his theory did not agree with all experimental data. The most important defects of Bohr's theory were the following-:

1. Bohr's atomic model was applied, mainly, on the hydrogen atom which is the simplest electronic system. It had succeeded in explaining the hydrogen spectral

lines only, however, it failed to explain the spectrum of any other element even that of the helium atom which contains only two electrons.

2. Bohr's theory considered the electron as a negative charged particle only, and' did not consider that it also has wave properties.

3. Bohr's theory postulated that it is possible to determine precisely both the location and speed of an electron at the same time. Indeed, this is experimentally impossible.

4. Bohr's theoretical equations described the electron as a particle moving in a circular planer orbit, which means that the hydrogen atom is planar. Later on it was confirmed that the hydrogen atom has three dimensional coordinates.

These aspects were sufficient to show the inadequacy of Bohr's theory. Scientists continued their research to understand the atomic structure as it actually exists in nature.

The Principles of Modern Atomic Theory

Modern atomic theory is based on some essential modifications on Bohr's model. The most important modifications are the following:

a) The dual nature of the electron.

b) The Heisenberg uncertainty principle.

c) Wave mechanical theory of the atom.

a) The dual nature of the electron :

All the previously mentioned theories considered the electron just as a minute negatively charged particle. However, the experimental data showed that the electron has a dual nature in the sense that it is a material particle which also has wave properties.

b) The Heisenberg uncertainty principle:

Bohr's theory postulated that it is possible to determine both the location and the velocity of the electron precisely at the same time. By applying the principles of quantum mechanics, Heisenberg concluded that "the determination of both the velocity and the position of an electron at the same time is practically impossible". We can only say that it is probable (to a greater locate or lesser extent) to locate the electron in this or that place. Thus, to speak in terms of "probability" seems to be more precise

Wave mechanical theory of the atom

In 1926, the Austrian scientist, Schrodinger applying the ideas of Planck, Einstein, De Broglie, and Heisenberg established the wave mechanical theory of the atom and managed to derive a wave equation that could describe the electron wave motion in the atom. On solving Schrodinger's equation it is possible to determine the allowed energy levels and to define the region of space around the nucleus where it is most probable to find the electron in each energy level. As a result of Schrodinger's work, our concept of the electronic motion around the nucleus has changed. Instead of speaking about the stable circular "orbits" of particular radii, and the areas between these orbits as being completely forbidden for electrons. The concept of electron cloud used to express the region of space around the nucleus where the possibility of finding the election in all distances, and directions. Inside the electron cloud there are areas that have a great possibility of finding an electron in it.

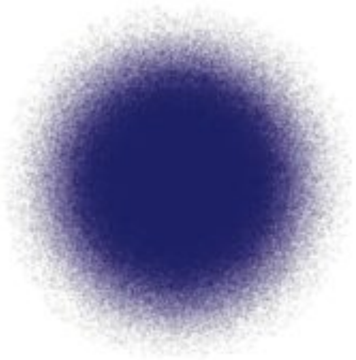


Fig: 1-7 Electron cloud

The mathematical solution of the Schrodinger equation introduced four numbers which are called quantum numbers .

Quantum numbers

To determine the energy of an electron in multielectrons atoms we should know the four quantum numbers which describe it, these four quantum numbers are :

1. The principal quantum number (n) describe the distance of the electron from the nucleus
2. The subsidiary quantum number (ℓ) describe the shape of electron cloud in sublevel
3. The magnetic quantum number (m_ℓ) describe the shape and number of orbital at which electrons present
4. The spin quantum number (m_s) describe spin motion of electron.

1- The principal quantum number (n) :

a) Bohr had used this number in explaining the spectrum of the hydrogen atom. It is given the symbol (n) and is used to define the following;

1. The order of the principal energy levels or electron shells. Their number in the heaviest known atom in its ground state is seven.

2. The number of electrons required to fill a given energy level equals two times the square of the shell number ($2n^2$),i.e.

The 1st shell K is filled with(2×1^2) =2 electrons.

The 2nd shell L is filled with (2×2^2) = 8 electrons.

The 3rd shell M is filled with (2×3^2) = 18 electrons.

The 4th shell N is filled with (2×4^2) = 32 electrons.

This rule does not apply to energy levels higher than the fourth level i.e. the fifth energy level should take theoretically 50 electrons, the 6th level takes 72 electrons, etc. However, the atom becomes unstable if the number of electrons exceeds 32 electrons on any level.

b) The principal quantum number is limited to any whole number value 1. 2, 3, 4,... etc, excluding zero. Each energy level is subdivided into a number of sublevels. Their energy is defined by the values of another quantum number called the subsidiary quantum number.

2- The subsidiary quantum number (ℓ):

a) It indicates the energy sublevels within each principal energy level.

b) Each principal energy level consists of a number of energy sublevels equal to its principal quantum number.

The energy sublevels take the symbols and values which are shown in the following table:

Symbols of sublevels	s	p	d	f
Values of subsidiary quantum number (l) [o: n-1]	0	1	2	3

It is observed that there is a small difference in the energy of the sublevels. They may be arranged according to increasing energy in the following order : s < p < d < f

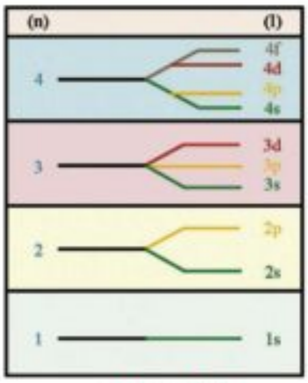


Fig: 1-8

3- The magnetic quantum number ($m\ell$):

The magnetic number is characterized by the following:

a) It represents the number of orbitals within a certain energy sublevel and their direction in space.

b) It is represented by odd and integer numbers between ($-\ell$,...,0,...$+\ell$). The following table explain probability of magnetic quantum number for atom (n=4).

(n)	(ℓ)	(ml)
1	0	0
2	0	0
	1	-1, 0, +1
3	0	0
	1	-1,0,+1
	2	-2,-1,0,+1,+2
4	0	0
	2	-2, -1, 0, +1, +2
	3	-3, -2, -1, 0, +1, +2, +3

c) Sublevel (s) has one orbital of spherical symmetrical shape around the nucleus. The sublevel (p) consists of three orbitals whose axes take the three spatial orientations (orientation in space x, y, z). Thus they are designated as px, py and pz . Each p orbital is perpendicular to the other two. The electron cloud of each orbital takes the form of two pears meeting head to head (dumb-bell shaped) at a node i.e. point of zero electron density.

See(1-9), sublevel (d) has 5 orbitals and sublevel (f) has 7 orbitals.

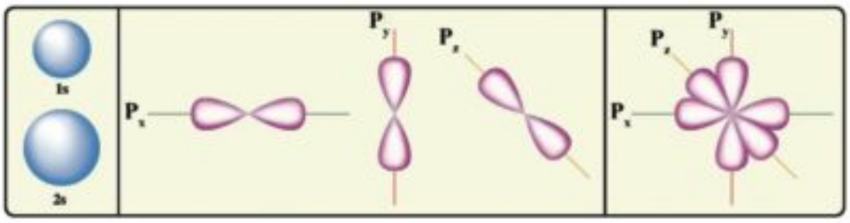


Fig: 1-9

4- The spin quantum number (m_s):

Any orbital cannot be occupied by more than two electrons, each electron spins on its own axis during its orbit around the nucleus. This can be illustrated when we imagine the spinning of the earth on its own axis during its rotation around the sun. Although the two electrons of the same orbital carry the same negative charge, we might expect them to repel. Yet as a result of the spinning of each electron on its own axis, a magnetic field arises in a direction opposing the direction of the other magnetic field arising from the spinning of the other electron. It is said that the two electrons are in a spin paired state and these are designated as (↑↓) See (1-10). The following considerations must be observed about the spin quantum number:

It defines the type of spin motion of the electron and since the electron can only spin in either of the two directions i.e. clock-wise (↑) $m_s = +\frac{1}{2}$ or anticlock-wise (↓) $m_s = -\frac{1}{2}$.

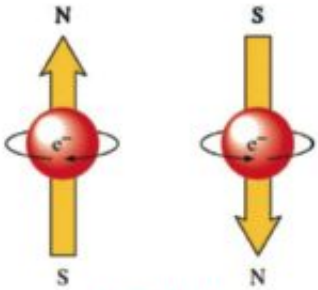


Fig: 1-10

An orbital electronic spin

Summary of the Relationship Between the Principal Energy level, Sublevels, and Orbitals

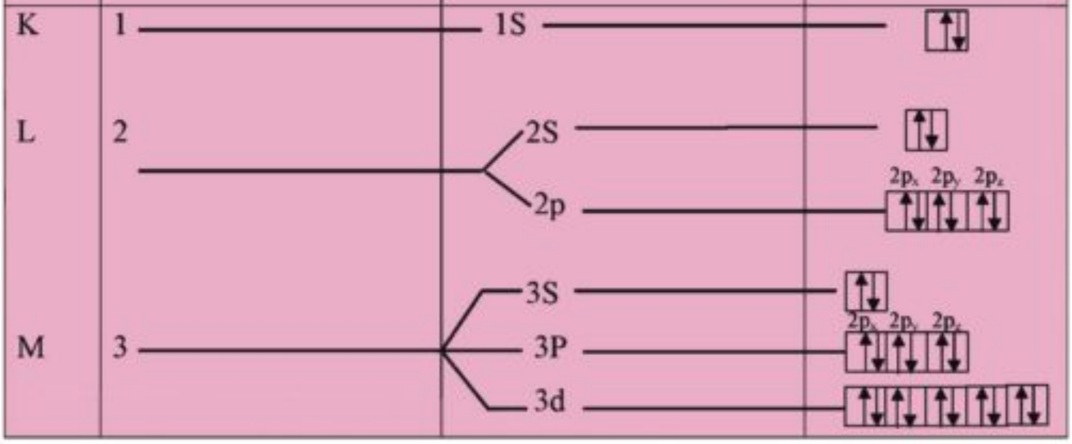
1. The number of energy sublevels equals the number of the principal level to which they belong, i.e. the first principal level consists of one sublevel and the 2nd principal level has two sublevels ...etc.

2. The number of orbitals within a principal energy level square the number of level (n^2), i.e. The 2nd energy level consists of 4 orbitals 2s, $2p_x$, $2p_y$, $2p_z$ and the 3rd energy level consists of nine orbitals (3S, $3p_x$, $3p_y$, $3p_z$ and five 3 d orbitals).

3. The number of electrons occupying a given principal energy level equals two times the square of this level ($2n^2$). For example, the 2nd level can take eight electrons distributed as follows $2s^2$, $2p_x^{\ 2}$, $2p_y^{\ 2}$, $2p_z^{\ 2}$.

The quantum numbers of the electrons occupying the first three energy levels may be summarized as follows:

Level	Principal quantum no.(n)	Subsidiary quantum no.(ℓ)	Magnetic q. no. $m\ell = 2\ell + 1$
	Define the energy levels	Define the no. of sublevels, which equal the principal quantum no.(n)	Define the no. of orbitals in each sublevel. S=1 ,P=3 d=5 ,f=7



Principles of distributing electrons

There are three important rules which must be considered in distributing electrons in the atom. These rules are :

1. Pauli's exclusion principle : it states that:

It is impossible for two electrons in the same atom to have the same four quantum numbers.

The following table explains two electrons of 3s similar in quantum numbers (n, ℓ, $m\ell$) but differ in (m_s):

4 quantum numbers	n	ℓ	$m\ell$	m_s
first electron	3	0	0	$+\frac{1}{2}$
second electron	3	0	0	$-\frac{1}{2}$

2. Aufbau (building-up) principle :

We have already seen that each energy level may consist of a number of energy sublevels which differ slightly from each other in energy, thus the real sequence of energy in the atom must follow the sequence of energy sublevels .

The Aufbau principle states that:

"Electrons occupy the sublevels in the order of increasing their energy, the lowest energy sublevels are filled first".

The sequence of energy sublevels according to their increasing energy follows the order :

1 S < 2 s < 2 p < 3 s < 3 p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5 f < 6d < 7p

Fig(1-11) illustrates a simple way to remember the order of filling of the energy sublevels by following the direction of the arrows.

4s "Energy"

$4s = n + \ell =$

$n+\ell = 4 + 0 = 4$

3d "Energy"

$3d = n+\ell =$

$3 + 2 = 5$

So, energy of

$4s < 3d$

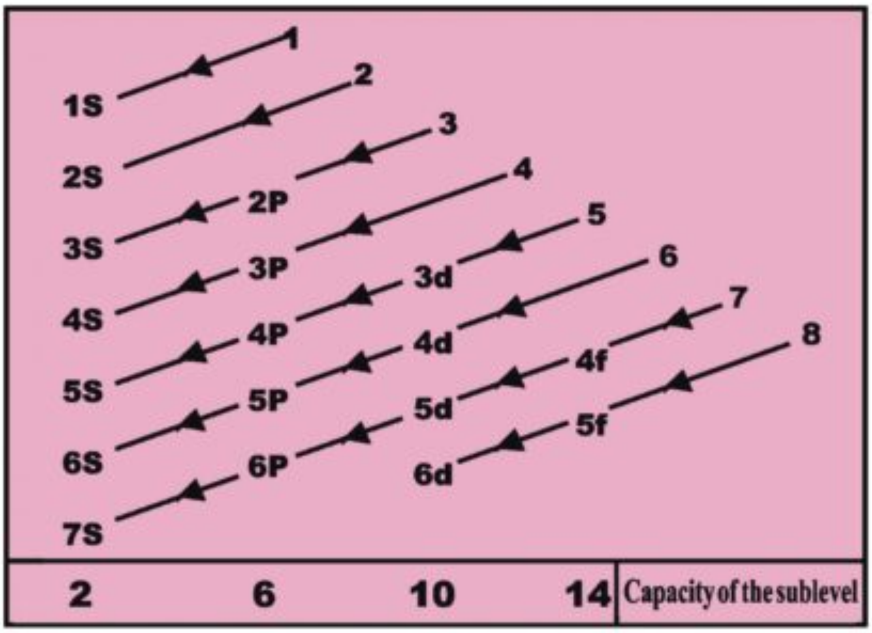


Fig: 1-11

Exercise :

Write the electron configuration for the following atoms :

$_{11}Na$, $_{20}Ca$, $_{30}Zn$ according to building up principle

Solution :

$_{11}Na$: $1s^2$, $2s^2$,$2p^6$, $3s^1$

$_{20}Ca$:$1s^2$, $2s^2$,$2p^6$, $3s^2$,$3p^6$,$4s^2$

$_{30}Zn$: $1s^2$, $2s^2$,$2p^6$, $3s^2$,$3p^6$,$4s^2$,$3d^{10}$

2. Hund's rule

states that:

" No electron pairing takes place in a given sublevel until each orbital contains one electron "

On writing the electron configuration of the nitrogen atom, (atomic no.7), we observed that the sublevel 2p has 3 electrons. Sublevel 2p consists of 3 orbitals which are equal in energy. How are these 3 electrons distributed in the 3 orbitals. According to Hund's rule, we find that each electron will occupy one orbital alone because this is more preferable from the energy point of view. When two electrons are paired in one orbital, inspite of their opposite spin, there must be a repulsion force which decreases the stability of the atom, i.e. increases its energy.

It should be noticed that the spin of single electrons must be in the same direction, because this gives the atom maximum stability.

In the oxygen atom, we find that sublevel 2p has 4 electrons, three of which are distributed first in the three orbitals according to Hund's; then the fourth electron has two possibilities, either it enters any orbital of those three 2p orbital and be paired with any electron of them where it will suffer a repulsion force with the electron already existing in the orbital, or it may enter the higher energy sublevel 3s. Obviously, it is still preferable from the energy point of view for an electron to be paired with another one in a lower energy level than to be alone in a higher energy level.

N.B. :

The following example explains the electronic configuration of nitrogen atom ($_7N$) according to Hund's rule.

* (Is^2 , $2s^2$, $2p_x^1$ $2p_y^1$ ، $2p_z^1$). This form explains the electron distribution in orbitals according to Hund's rule.

Evaluation

QUESTION I:

Choose the correct answer:

1- **The first scientist who defined the element is.............**

a. Boyle b. Rutherford c. Bohr d. Thomson.

2- **All matters are composed of four components (water, air, dust, and fire) with a different ratio. That idea belongs to ....**

a. Bohr b. Rutherford c. Dalton d. Aristotle.

3- **The strong evidence that proved that cathode rays exist in all matters .....**

a. they have thermal effect.

b. flow in straight lines.

c. consist of fine particles.

d. they have the same properties and behavior whatever the gas or the cathode material used.

4- **Cathode rays consists of particles called ....**

a. alpha particle

b- electrons

c - atoms

d - orbitals

5- **The uncertainty principle was found by............**

a. Schrodinger b. De Broglie c. Rutherford d. Heisenberg

6- s, p, d, and f are symbols of ....

a. principal energy levels.

b. energy sublevels.

c. orbitals of the sublevel.

d. the single electrons in the sublevel.

7. The quantum number which define the spin motion of the electron is...........

a. the principal quantum number

b. the subsidiary quantum number

c. the magnetic quantum number

d. the spin quantum number

8- **Which configuration represents nitrogen according to Aufbau ....**

a. 2 ,5 b. $Is^2, 2s^2, 2p^3$ c. $Is^2, 2s^2, 2p_x^1 2p_y^1, 2p_z^1$ d. $Is^2, 2s^1, 2p^4$.

9- **Heating gases or vapours to high temperature under reduced pressure, they..........**

a. absorb energy

b. emit light

c. emit gamma radiation

d. emit Alpha radiation

10 - When an electron absorb a quantum of energy it ....

a. Transfers to all higher energy levels

b. Transfers to the higher energy level corresponding to the absorbed energy

c. Transfers to lower energy level

d. Transfers to the lower energy level that corresponds to the absorbed quantum

11-**The magnetic quantum number (m/) defines ....**

a.The principal energy level.

b.The number of energy sublevels.

c.The number of orbitals and their shape.

d. The number of electrons in orbitals.

12- The no. of orbitals in the sublevel 3d equal ....

a. 5 b. 4 c. 6 d. 7

13- The no. of orbitals in the principal energy level (n) equals ....

a. $2n^2$ b. $3n^2$ c. n^2 d. (n-1)

14- The maximum no. of electrons that occupy a given energy level (n) equals............

a. 2n b. n^2 c. $2n^2$ d. $(2n)^2$

15- The energy sublevel may be arranged according to their increasing energy in an ascending order as following:

a. 3s< 3p<4d< 4s b. 3s< 4p < 3d < 4f

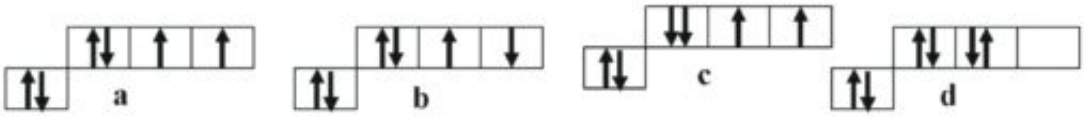
c. 3s< 3p < 3d < 4s d. 3s<3p <4s<3d

16- The orbitals of the same energy sublevel are ....

a. different in energy b. equal in energy

c. different in shape d. (a and c) correct

17 - One of the following diagrams shows the correct distribution of electrons in the last energy level of oxygen atom .



QUESTION 2:-

a- Explain the observations upon which Rutherford reached the following conclusions:

a. Most of the atom is an empty space and it is not a solid sphere.

b. There is a very dense tiny piece of the atom later on named the nucleus.

c. The charge of the dense part of the atom in which most of its mass is concentration should have a positive charge similar to alpha particles.

QUESTION 3:-

Explain how the cathode rays may be obtained.

QUESTION 4:-

Explain Thomson's atomic model.

QUESTION 5:-

Write the probability of four quantum numbers of the last electron for the following elements

B_5 Boron   Fluorine ($_9F$)   Sodium ($_{11}Na$)

QUESTION 6:-

What are possible values of (ℓ) when (n=3) ?

QUESTION 7:-

What is meant by

a. Electron cloud

b. The dual nature of electron

c. The building up principle

d. Hund's rule

e. The Heisenberg uncertainty principle

f. Pauli exclusion principle

QUESTION 8:-

Write the electronic distribution of the following atoms :

$_{35}Br$, $_{30}Zn$, $_{26}Fe$, $_{18}Ar$, $_{20}Ca$, $_{11}Na$

QUESTION 9:-

Give reason for :

a. The line spectra of any element is a specific.

b. The electron has a dual nature.

c. The atom is electrically neutral.

d. No electron pairing takes place until each orbital contains one electron.

c. The sublevel p takes up to 6 electrons whereas sublevel d takes 10 electrons.

QUESTIQN 10:-

Which of the following quantum numbers for an electron include wrong? giving reason ?

a) $n = 3$, $\ell = 2$, $m\ell = -1$, $m_s = +\frac{1}{2}$
b) $n = 4$, $\ell = 3$, $m\ell = -2$, $m_s = +\frac{1}{2}$
c) $n = 1$, $\ell = 1$, $m\ell = 1$, $m_s = -\frac{1}{2}$

QUESTION 11:-

Write the possible values (ℓ) , (ml) for the electron its principle quantum number, ($n = 2$).

Chapter Two

The periodic table and classification of elements

Objectives

At the end of this chapter, the student should be able to:

- Describe the long form periodic table.
- Arrange the energy sublevels according to the building up principle.
- Identify the type of the element and its properties from its location in the table.
- Calculate the atomic radius by using Bond length.
- Explain the factors affecting the atomic radius across the periods and groups.
- Define the location of the four blocks of the table.
- Find the relationship between the electronic configuration of the elements of the same groups.
- Define the atomic radius, ionization energy, electron affinity, and electronegativity.
- Compare between the electron affinity and electronegativity.
- Identify the location of metals and nonmetals.
- Find the relationship between atomic radius, ionization energy, and electron affinity in metals and nonmetals.
- Identify the relationship between the atomic radius and the acidic and basic properties.
- Discuss the ionization of acids and bases as hydroxyl compounds.
- Calculate the oxidation number.
- Explain the oxidation and reduction in different reaction.

The long form periodic table

As more knowledge about atomic structure was gained, and as the actual energy levels in the atom (i.e. energy sublevels) were discovered and as the building up (auf bau) principle was reached, The long form periodic table was constructed on the bases of :

1 - The sequence of elements agrees with the auf bau principle i.e. the sequence of filling the atomic energy sublevels.

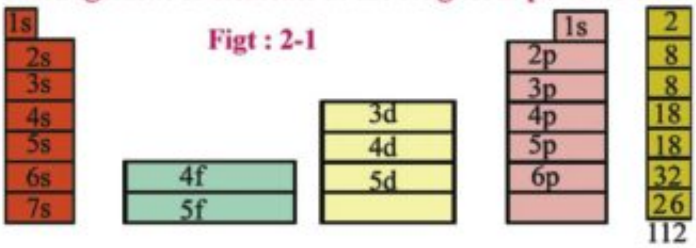
2 - Elements are arranged so that each element has one electron more than the element before it. Thus if we recall the sequence of atomic sublevels according to increasing energy, we observe that they agree well with the sequence of elements in the long from periodic table as follows:

Sequence of energy sublevels:

1s | 2s < 2p | 3s < 3p | 4s < 3d < 4p

5s < 4d < 5p | 6s < 4f < 5d < 6p | 7s < 5f < 6d

Arrangement of elements in the long form periodic table



Figt : 2-1

The table may be divided into four main zones or blocks these blocks are :

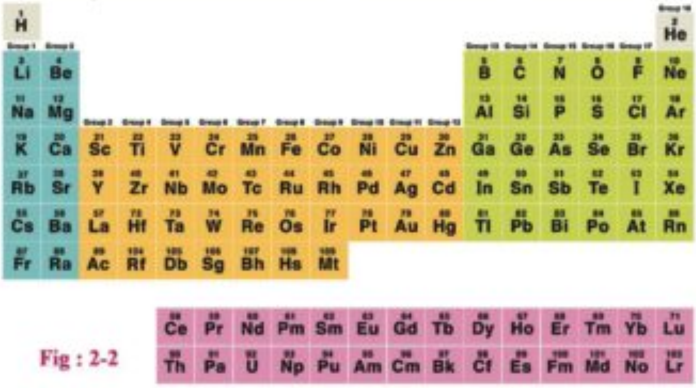


Fig : 2-2

a) S-Block elements:

They are placed in the left hand block of the table. The outermost electrons of this group of elements occupy the s-sublevel.The s-block consist of two groups of elements. The first is group I-A, whose elements have the configuration ns^1, and the second is group II-A, whose elements have the configuration ns^2, where n is the number of the outer energy level and the number of period at the same time.

b) p-Block elements:

They occupy the right hand block of the table. The p-block contains the elements whose outermost electrons occupy the p-Sublevel.They are six groups (i.e. III-A ,IV-A , V-A , VI-A , VII-A and group number zero of noble gases). The configuration of p-block elements is np^1 in group (III), np^2 in group (IV), np^3 in group (V) ... and so on, until is becomes completely filled np^6 in group zero. The elements of the groups of s and p-blocks are known as the representative or main group elements i.e. they include all the A-groups.

c) d-Block elements:

They occupy the middle block of the table which contains the elements with the outermost electrons occupying the d sublevel. Since the d-sublevel can take up to ten electrons, it contains ten vertical columns, seven of which belong to B-groups and the other three of them belong to group (VIII).

The d-block elements are known as transition elements and are themselves classified according to the number of the outer energy level and according to the period, giving three series which are :

1. The first transition series :

It includes the elements in which the sublevel 3 d is filled successively.

They are placed in the fourth period and consist of the elements from scandium to zinc.

2. The second transition series :

It includes the elements in which the sublevel 4d is filled successively. They are placed in the fifth period and consist of the elements from yttrium to cadmium.

3. The third transition series :

It includes the elements in which the sublevel 5 d is filled successively. They are placed in the sixth period and consist of the elements from lanthanum to mercury.

d) f -Block elements :

This includes the elements in which the f-sublevel is filled successively, Sublevel f can take up to 14 electrons. The f-block includes two series i.e. the lanthanide and the actinide series.

1. The lanthanide series :

In this series the sublevel 4f is filled successively so it consists of 14 elements. It must be taken into consideration that the outermost energy level for all these elements are $6s^2$, so they are quite similar in behavior.

Consequently, they are very difficult to be separated and that is why they are known as rare earths.

2. The actinide series :

In this series the sublevel 5 f is filled successively so it includes 14 elements. All the actinides are radioactive elements and their nuclei are unstable. The f-block elements are known as the inner transition elements. They are usually separated from the table (placed below it). So that it is not too wide.

This shows the advantage of this table, that it can be separated into blocks. It is clear now that it is possible to classify the elements in the long form periodic table into four types.

1. Noble gases:

They are the elements of the last column of the p-block (group zero). Their electronic structure is np^6 except for Helium which is $1s^2$. They are characterized by having energy levels completely filled by electrons. Consequently, they are very stable elements, they may form compounds but with great difficulty.

2. The representative elements:

They are elements of s and p-block except that of group zero. These elements are characterized by the complete filling of all the energy levels with electrons except for the highest level. Their highest level tends to reach the completed configuration ns^2, np^6 by gaining, losing, or sharing electrons i.e.

3. The main transition elements :

They are the elements of the d-block. The d-sublevels of these elements are successively filled.

4. The inner transition elements :

These are the elements of the f-block. The f-sublevels of these elements are successively filled.

Modern periodic table

Key: 6 — Atomic number; C — Symbol; Carbon — Name

s block elements: groups IA (1), IIA (2) · d block elements: groups IIIB (3) to IIB (12) · p block elements: groups IIIA (13) to 0 (18)

Period	IA (1)	IIA (2)	IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	VIII (8)	VIII (9)	VIII (10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	0 (18)
First period	1 H Hydrogen																	2 He Helium
Second period	3 Li Lithium	4 Be Beryllium											5 B Boron	6 C Carbon	7 N Nitrogen	8 O Oxygen	9 F Fluorine	10 Ne Neon
Third period	11 Na Sodium	12 Mg Magnesium											13 Al Aluminum	14 Si Silicon	15 P Phosphor	16 S Sulfur	17 Cl Chlorine	18 Ar Argon
Fourth period	19 K Potassium	20 Ca Calcium	21 Sc Scandium	22 Ti Titanium	23 V Vanadium	24 Cr Chromium	25 Mn Manganese	26 Fe Iron	27 Co Cobalt	28 Ni Nickel	29 Cu Copper	30 Zn Zinc	31 Ga Gallium	32 Ge Germanium	33 As Arsenic	34 Se Selenium	35 Br Bromine	36 Kr Krypton
Fifth period	37 Rb Rubidium	38 Sr Strontium	39 Y Yttrium	40 Zr Zirconium	41 Nb Niobium	42 Mo Molybdenum	43 Tc Technetium	44 Ru Ruthenium	45 Rh Rhodium	46 Pd Palladium	47 Ag Silver	48 Cd Cadmium	49 In Indium	50 Sn Tin	51 Sb Antimony	52 Te Tellurium	53 I Iodine	54 Xe Xenon
Sixth period	55 Cs Cesium	56 Ba Barium	57 La Lanthanum	72 Hf Hafnium	73 Ta Tantalum	74 W Tungsten	75 Re Rhenium	76 Os Osmium	77 Ir Iridium	78 Pt Platinum	79 Au Gold	80 Hg Mercury	81 Tl Thallium	82 Pb Lead	83 Bi Bismuth	84 Po Polonium	85 At Astatine	86 Rn Radon
Seventh period	87 Fr Francium	88 Ra Radium	89 Ac Actinium	104 Rf Rutherfordium	105 Db Dubnium	106 Sg Seaborgium	107 Bh Bohrium	108 Hs Hassium	109 Mt Meitnerium	110 Ds Darmstadtium	111 Rg Roentgenium	112 Uub Ununbium	113 Uut	114 Uuq	115 Uup	116 Uuh	117 Uus	118 Uuo

f block elements

Lanthanides	58 Ce Cerium	59 Pr Praseodymium	60 Nd Neodymium	61 Pm Promethium	62 Sm Samarium	63 Eu Europium	64 Gd Gadolinium	65 Tb Terbium	66 Dy Dysprosium	67 Ho Holmium	68 Er Erbium	69 Tm Thulium	70 Yb Ytterbium	71 Lu Lutetium
Actinides	90 Th Thorium	91 Pa Protactinium	92 U Uranium	93 Np Neptunium	94 Pu Plutonium	95 Am Americium	96 Cm Curium	97 Bk Berkelium	98 Cf Californium	99 Es Einsteinium	100 Fm Fermium	101 Md Mendelevium	102 No Nobelium	103 Lr Lawrencium

Description of the long form periodic table

- It consists of 18 vertical groups (columns) and seven horizontal periods.

-The elements are arranged ascendingly according to their increasing atomic numbers.

- Each element in the same period increases by one electron. Every new period begins by filling a new energy level of larger principal quantum number with one electron. Then by successive filling of atomic energy sublevels lying in the same period, we reach the last element in the period which is the noble gas. The elements of the same vertical group are identical in the electron composition of the highest energy level except that the principal quantum number of the highest level change by one down the group.

The following table explains the electronic configuration of the first thirty elements in the periodic table.

Electronic configuration of the first thirty elements in the periodic table

		1	2		3			4				5				6				7
Z	Element	s	s	p	s	p	d	s	p	d	f	s	p	d	f	s	p	d	f	s
1	H	1																		
2	He	2																		
3	Li	2	1																	
4	Be	2	2																	
5	B	2	2	1																
6	C	2	2	2																
7	N	2	2	3																
8	O	2	2	4																
9	F	2	2	5																
10	Ne	2	2	6																
11	Na	2	2	6	1															
12	Mg	2	2	6	2															
13	Al	2	2	6	2	1														
14	Si	2	2	6	2	2														
15	P	2	2	6	2	3														
16	S	2	2	6	2	4														
17	Cl	2	2	6	2	5														
18	Ar	2	2	6	2	6														
19	K	2	2	6	2	6		1												
20	Ca	2	2	6	2	6		2												
21	Sc	2	2	6	2	6	1	2												
22	Ti	2	2	6	2	6	2	2												
23	V	2	2	6	2	6	3	2												
24	Cr	2	2	6	2	6	5	1												
25	Mn	2	2	6	2	6	5	2												
26	Fe	2	2	6	2	6	6	2												
27	Co	2	2	6	2	6	7	2												
28	Ni	2	2	6	2	6	8	2												
29	Cu	2	2	6	2	6	10	1												
30	Zn	2	2	6	2	6	10	2												

Trends and periodicity of properties in the periodic table:

We have explained the principles of arranging the elements in the periodic table and have discussed the relations between the electron structure of the element and its location in the table. We will now explain the graduation of physical and chemical properties in periods and their relation with the electron structure .We will concentrate our study here on the trends in properties of the main group elements only, *i.e.* the s and p block elements. The study of the trends in the properties of transition elements will be considered elsewhere.

1. The atomic radius :

The wave mechanics theory reveals that it is impossible to determine the precise location of an electron around the nucleus, consequently, the atomic radius cannot be defined and cannot be physically measured. Thus, it is incorrect to define the atomic radius as the distance from the nucleus to the farthest electron. Instead, the atomic radius is defined as :

"Half the distance between the centers of two similar atoms in a diatomic molecule".

The distance between the nuclei of two bonded atoms is known as the bond length.

Molecule	Bond length by A°	Covalent atomic radius
H - H	0.60	0.30
F - F	1.28	0.64
Cl - Cℓ	1.98	0.99
Br - Br	2.28	1.14
I - I	2.66	1. 33

Exercise :

The bond length in the chlorine molecule Cℓ-Cℓ is 1.98 A° and the bond length between carbon and chlorine atoms C-Cℓ = 1.76 A°. Calculate the atomic radium of carbon.

Solution :

The atomic radium of chlorine (Cℓ) atom

$$\frac{\text{bond length}}{2} = \text{in chlorine molecule} = \frac{1.98}{2} = 0.99A^o$$

The atomic radius of carbon = bond length between carbon and chlorine - atomic radius of chlorine = 1.76 - 0.99 = 0.77A°

In ionic crystals constructed of cations and anions (e.g the sodium chloride crystal), the distance between the centers of cation and anion in the formula unit can be measured by the technique mentioned previously. The distance between the centers of two atoms is the sum of their two radii. The atomic radius measured in this way is known as the ionic radius and depends on the number of electrons lost or gained to form the ions.

Fig 2-3 represents the graduation of atomic radius in nontransition elements

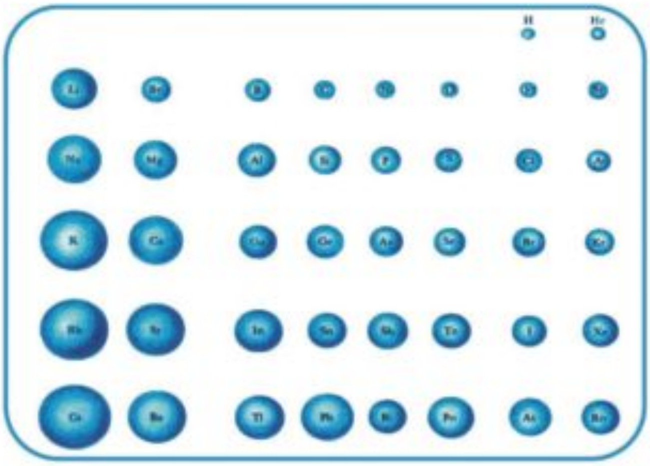


Fig 2:3

a) In the horizontal periods :

We see that the atomic radii decrease as we go from the left to right across a period in the periodic Table (i.e, if we start from the first group and go to group zero). This is due to the increase in the nuclear charge gradually (effective nuclear charge) (Z-effect) which known by, actual nuclear charge that affects an electron in an atom,The effective nuclear charge always less than the nuclear charge (Number of protons) as a result of screening effect of the inner electrons (core electrons) on part of this change of the electron we study. which increasingly attracts the valence electrons leading to decrease the atomic radius. This means that the biggest atom in size in a given period is that of the first group, and the smallest atom is that of group seven (halogens).

b) In the vertical groups:

On descending vertically a group in the periodic table, the atomic radii increase with the increase in atomic number. This is attributed to:

1. The effect of extra shells of electrons being added.
2. The inner filled atomic orbitals having a screening effect on the pull of the nuclear charge on the outer electrons.
3. The increased repulsive forces between electrons.

Exercise :

Explain the variation in atomic and ionic radii for sodium, iron and chlorine as shown in the following table.

atom or ion	Na	Na^+	Fe	Fe^{2+}	Fe^{3+}	$C\ell$	$C\ell^-$
radius A	1.75	0.95	1.17	0.75	0.60	0.99	1.81
no. of protons	11	11	26	26	26	17	17
no. electrons	11	10	26	24	23	17	18

Solution :

In the case of metals like sodium and iron, it is observed that the radius of cation is smaller than that of the atom. This is due to the increasing pull of the effective nuclear charge on the remaining electrons in the cations. As the ionic charge increases as from iron (II) to iron (III) the radius decreases. However, in the case of nonmetals, like chlorine, the anionic radius is larger than the atomic radius due to the increase in the number of electrons without increasing the nuclear charge (Have you played tug of war. It is very much the same.)

The trends in other physical properties like the melting and boiling points and density will be studied later on.

2. Ionization potential (ionization energy) :

If energy is supplied to an atom, electrons may be excited and transferred to higher energy levels. But if sufficient energy is supplied the most loosely bound electrons may be completely removed, giving a positive ion. The ionization energy is defined as follows :

Ionization energy:

"It is the amount of energy required to remove the most loosely bound electron completely from an isolated gaseous atom".

Since it is possible to remove one, two, Three or more electrons from most atoms there are first, second, third ... etc. ionization energies,

The first ionization energy :

This is the energy required to convert an atom to an ion with one positive charge. $M \longrightarrow M^{+} + e^{-}$

The second ionization energy :

it is the energy required to convert M^{+} to M^{2+}

$M^{+} \longrightarrow M^{2+} + e^{-}$

The following trends have to be considered on the first ionization potential

a) In periods :

In general the first ionization energy increases on crossing a period as we move from left to right due to increasing effective nuclear charge and also the atomic radius decreases. Which leads to increasing nuclear attraction force to valency election which need higher energy to separate them from the atom. This means that the ionization energy is inversely proportional to the atomic radius.

b) In groups :

The ionization energy decreases in descending groups, in accordance with the increasing of atomic number, because in descending a group extra shells of electrons are added which increase the atomic size and so nuclear attraction force for valency electrons decreases and the energy required to remove it, becomes smaller .

c) The first ionization energy of noble gases of group zero is very high.

This is due to the stability of their electronic configuration, because it is difficult to remove an electron from a completely filled shell.

d) The second ionization energy is greater than the first one due to the increased effective nuclear charge. The third ionization energy is much greater because it results from the breaking up of a completely filled shell. This is shown in the ionization energies of magnesium.

$$Mg(2,8,2) \longrightarrow Mg^{+} + e^{-}, \Delta H = +737 \ kJmol^{-1}$$

$$Mg^{+}(2,8,1) \longrightarrow Mg^{2+} + e^{-}, \Delta H = +1450 \ kJ\ mol^{1-}$$

$$Mg^{2+}(2,8) \longrightarrow Mg^{3+} + e^{-}, \Delta H = +7730 \ kJ\ mol^{-1}$$

3. Electron affinity :

We have mentioned that the removal of an electron from the atom will convert it a cation, which require energy (the first ionization energy). On the other hand if the atom gained an extra electron, it will be converted to anegative ion. This is associated with release of energy. Electron affinity may be defined as follows :

> **"This is the amount of energy released when an extra electron is added to a neutral gaseous atom".**

This may be illustrated by the equation :

$$X + e^{-} \longrightarrow X^{-} + energy \qquad X + e^{-} \longrightarrow X^{-} \ , (\Delta H = -)$$

The graduation of electron affinity in the periodic table may be explained as follows:

1- The electron affinity decreases across the elements of same group by

increasing the atomic number . This is due to the increase of the atomic radius so the pull of the nucleus to the electrons decrease.

2- The electron affinity increases across the periods as we move from left to right i.e. by increasing atomic number. This is due to the gradual decrease in atomic size which makes it easier for the nucleus to attract the new electron and the values of electron affinity of neon is , berillium and nitrogen nearly zero. The exception in the case of beryllium is due to the stability.

of its atom, since it has filled orbitals (Is^2 , $2s^2$). In nitrogen atom the sublevel 2p has three electrons i.e. it is half-filled. The half-filled orbitals give the atom some extra stability. In neon all the sublevels are filled with electrons, which gives the atom great stability.

3- The magnitude of the electron affinity is high when the added electrons make the orbitals, half-filled or completely filled since in both cases this helps the stability of the atom.

4- It is observed that the electron affinity of fluorine is (-328 kJ / mol) which is less than that for chlorine (-348.6 kJ / mol). Since the fluorine atom is smaller in size it should be expected that its electron affinity is bigger. The smaller value for fluorine atom is attributed to the very small size of its atom. Thus the entering electron will suffer a strong repulsion force with the nine electrons already existing around the nucleus.

4- Electronegativity :

The electronegativity is defined as follows :

Electronegativity " it is the tendency of an atom to attract the electrons of the chemical bond to itself"

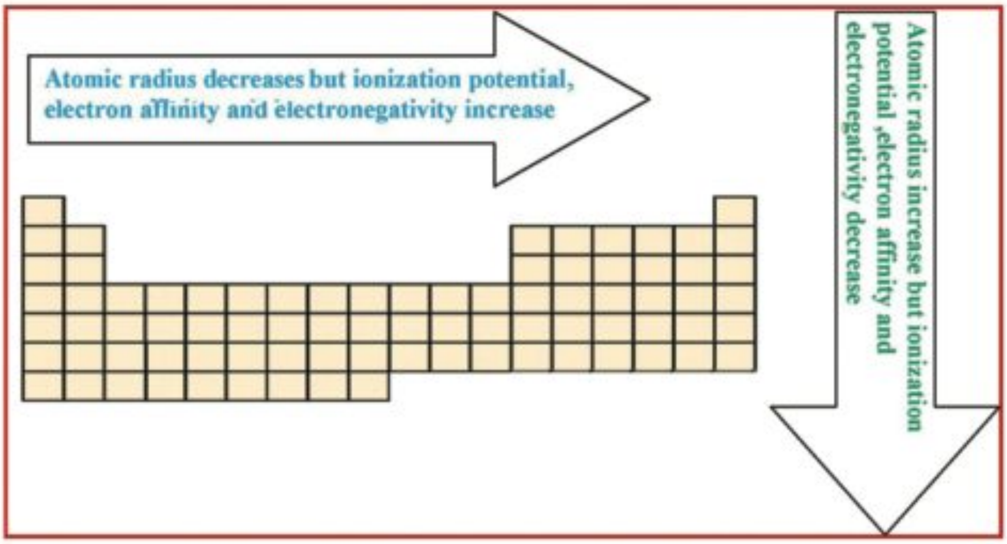
We must differentiate between electronegativity and electron affinity.

The latter term is an energy term which refers to the atom in its single state, while electronegativity refers to the combined atom.

Electronegativity for the element is expressed by values. A high value indicates increasing its relative ability to attract bond electrons a high electronegativity.

Electronegativity increases across the periods with the increase of the atomic number and decrease in the atomic radius in groups, however, electronegativity decreases with an increase of the atomic number.

Fluorine is considered to be the most electronegative element. It should be noticed that the difference in electronegativity between elements plays a very important role in determining the nature of the bond formed between them, as will be discussed later in chapter three.



5. Metallic and nonmetallic property:

We have mentioned that at the beginning of the nineteenth century (**Berzelius**) was the first scientist to classify elements into two main groups, i.e. metals and nonmetals. Indeed that was before anything was known about atomic structure. Although, this is an old classification, it is still currently in use .

Boundaries between them in the periodic table and their properties are blurred. With the development of our concept about the electron structure of atoms we can differentiate between metals and nonmetals as follows :

Metals :

1- This is the group of elements whose valence shell generally has less than half its capacity of electrons.

2- They lose electrons of the valency shell to reach the structure of the Nobel gas;(the aim of chemical reaction). Consequently, metals form positive ions and they are described as electropositive elements.

3- Their good electric conductivity is attributed to the mobility of their few valence electrons. These can transfer from one position to another in the metal structure.

4- Metals are characterized by their large atomic radius which leads to small magnitudes for ionization energy and electron affinity.

Nonmetals :

1- This is the group of elements whose valency shell generally has more than half its capacity of electrons.

2- To form ionic compound when reacting with metals they gain a small number of electrons to reach the structure of the noble gas and become negative ions. Thus nonmetals are described as electronegative elements.

3- They do not conduct electricity because their valence electrons are strongly bound to the nucleus. Thus, it is difficult for these valence electrons to be transferred. Therefore, nonmetals are insulators and non-conducting electricity.

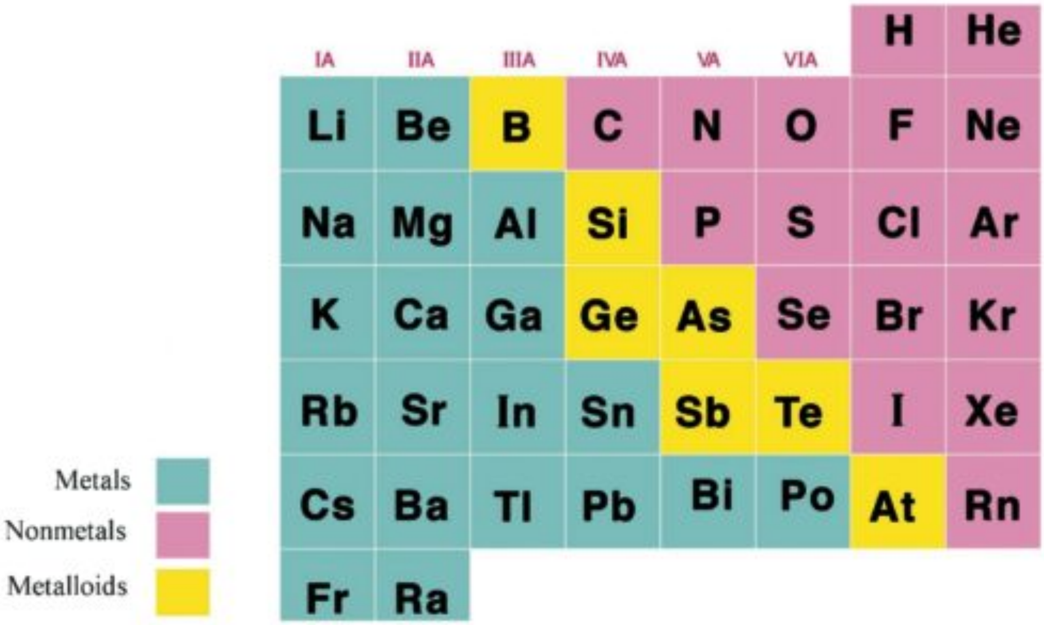
4- The small atomic radius of nonmetals leads to their high values of ionization energy and electron affinity.

There is a third group of elements that has a metallic appearance and has most of the properties of nonmetals at the same time. These elements are known as metalloids. They are characterized by the following properties:

1. Electronegativiy is intermediate between metals and nonmetals.

2. Their electrical conductivity is less than that of metals, but more than that of nonmetals.

3. They are used as semiconductors and are known as transistors.



It is clear from the above figure that metals are placed to the left of metalloids while nonmetals are placed to the right of metalloids. On the basis of ionization energy and electron affinity of elements, we can define metallic and non-metallic character trend in the periodic table as follows;

a) In periods :

As we move across the period from left to right we observe that the first group includes the elements of the highest metallic character. Then this properly decreases gradually with the increase in the atomic number across

the period past the metalloids. To the right of the metalloids begins the nonmetallic character. Group seven includes the elements of the highest nonmetallic character,

b) In groups :

The metallic character increases with the increase in the atomic number in descending group. Consequently, we conclude that the elements of strongest metallic character are placed at the bottom on the left hand side of the table. Thus caesium is considered as the element which has the highest metallic character.

On the other hand, the elements which the highest nonmetallic character is found at the top of the right side of the table. Thus fluorine is considered as the element which has the highest nonmetallic character.

6. Acidic and basic properties :

It is known that on dissolving nonmetal oxides in water they form acids.

$$\underset{(g)}{CO_2} + \underset{(\ell)}{H_2O} \longrightarrow \underset{(aq)}{H_2CO_3} \text{ carbonic acid}$$

$$\underset{(g)}{SO_3} + \underset{(\ell)}{H_2O} \longrightarrow \underset{(aq)}{H_2SO_4} \text{ sulphuric acid}$$

Thus, nonmetallic Oxides are usually known as acidic oxides. They react with alkalies forming salts and water.

$$\underset{(g)}{CO_2} + \underset{(aq)}{2NaOH} \longrightarrow \underset{(aq)}{Na_2CO_3} + \underset{(\ell)}{H_2O}$$

On the other hand, metallic oxides are usually known as basic oxides- Some basic oxides are soluble in water, others are not. The water soluble basic oxides are also known as alkalies.

$$\underset{(s)}{Na_2O} + \underset{(\ell)}{H_2O} \longrightarrow \underset{(aq)}{2NaOH}$$

$$\underset{(s)}{K_2O} + \underset{(\ell)}{H_2O} \longrightarrow \underset{(aq)}{2KOH}$$

Basic oxides react with acids forming salts and water.

$$\underset{(s)}{Na_2O} + \underset{(aq)}{2HC\ell} \longrightarrow \underset{(aq)}{2NaC\ell} + \underset{(\ell)}{H_2O}$$

$$\underset{(s)}{MgO} + \underset{(aq)}{H_2SO_4} \longrightarrow \underset{(aq)}{MgSO_4} + \underset{(\ell)}{H_2O}$$

There is a third type of oxides known as amphoteric, like aluminium oxide Al_2O_3, zinc oxide ZnO, antimony oxide Sb_2O_3 and tin oxide SnO. These oxides react either as basic oxides or as acidic oxides.

$$\underset{(s)}{ZnO} + \underset{(aq)}{H_2SO_4} \longrightarrow \underset{(aq)}{ZnSO_4} + \underset{(\ell)}{H_2O}$$

$$\underset{(s)}{ZnO} + \underset{(aq)}{2NaOH} \longrightarrow \underset{\substack{(aq)\\ \text{sodium zincate}}}{Na_2ZnO_2} + \underset{(\ell)}{H_2O}$$

a) In periods :

We observe the acidic character in oxides increases when the atomic number increases and the basic character decreases.

b) In groups :

If we consider the elements of the first group as example to explain the trends in the basic property . We find that it increases in descending the groups or with increasing atomic number . This is due to the increases in the atomic size of the element, while the charge remains constant. Considering acids and bases as hydroxy compounds, they may be represented by the general formula (MOH), (where M is the element atom), It may be ionized by either ways :

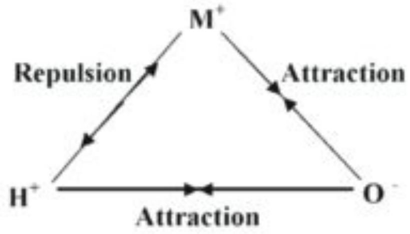
1- It may produce hydroxide ions and considered a base

$MOH_{(Base)} \rightleftharpoons M^+ + OH$

2- Or it may produce hydrogen ions and considered an acid

$MOH_{(Acid)} \rightleftharpoons MO + H^+$

Assume the three atoms of MOH are arranged in a triangle as follows :



a. If the attraction force between $M^+ O^-$ is bigger than that between H^+ , O^- the substance will ionize as an acid.

b. If the attraction force between H^+ and O^- is bigger than that between M^+ and O^- , the substance will ionize as a base.

c. If the attraction forces are equal, the substance will ionize as an acid or a base depending on the reaction medium this means that it reacts as base in the acidic medium and as an acid in the basic medium.

The attraction forces in the previous reactions depends on the volume and the charge of the M atom . In alkali metals like sodium, we observe that sodium atom has a big volume however, it has only one positive charge.

Accordingly the attraction between Na^+ and O^- is weaker than the attraction between O^- and H^+ so OH^- ion is produced i.e it is ionized as a base . However, if we move in the same period to the right, we observe that the nonmetal atoms as chlorine has a small volume and a big charge which increases its attraction to O^- and ionized as an acid . The strength of oxygenated acids depends on the number of oxygen atoms which are not linked to hydrogen atoms.

If we represent the oxygenated acid by the formula $[MO_n(OH)m]$, where M is the element atom , we observe that the strong acid is that which has more number of non bonded oxygen atoms (O_n) with hydrogen.

Acid $MO_n(OH)_m$	No.of nonbinded Oxygen atoms with hydrogen	Strength of the acid
Orthosiliconic H_4SiO_4 $Si(OH)_4$	Zero	Weak
OrthophosphoricH_3PO_4 ' $PO(OH)_3$	1	Moderate
Sulphuric H_2SO_4 $SO_2(OH)_2$	2	Strong
Perchloric $HC\ell O_4$ $C\ell O_3(OH)$	3	Very strong acid

7. Oxidation numbers: In modern chemistry we use the term oxidation number instead of the term valency.

Oxidation number "It is a number that refers to the electric charge (positive or negative) that the atom or ion would have in the compound , be it ionic or covalent"

Rules for assigning oxidation number :

1. In ionic compounds :

The oxidation number of any atom is equal to its valency preceded by a positive sign in case of cations and with a negative one in case of anion. If the oxidation number is positive, this indicates that the number of electrons that the atom has lost to give this.

If the oxidation number is negative this indicates that the number of the electrons that the atom has gained to give this.

For example : K^+Br^-, $Na^+C\ell^-$, $Mg^{2+}O^{2-}$, $Ca^{2+}(CO_3)^{2-}$, $Cu^{2+}(SO_4)^{-2}$

2- In covalent compounds :

Since there are no negative or positive ions, we consider the charge carried on the atom explains the electronic shift in the chemical bond. The more electronegative atom carries a negative charge and the less electronegative one carries a positive charge. There are two cases in assigning the oxidation number in covalent compounds they are :

a. In molecules of similar atoms e.g. $C\ell_2$,O_3,P_4,S_8 . the electronic shift in the bonds between the atoms are equal, because the electronegativity of the atoms forming the molecule are similar, Accordingly, the oxidation number of any atom in this molecule is zero. fig (2-5)

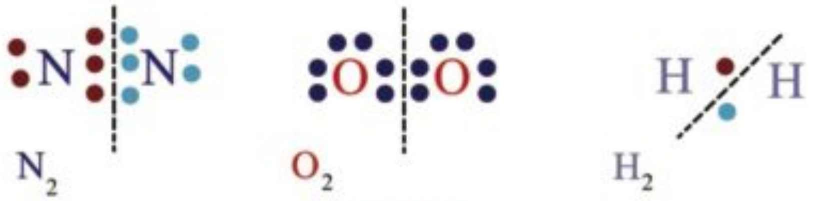


Fig: 2-5

b) In diatomic molecules of different atoms in electronegativity, the shared electrons are assigned to the more negative atom. Fig (2-6)

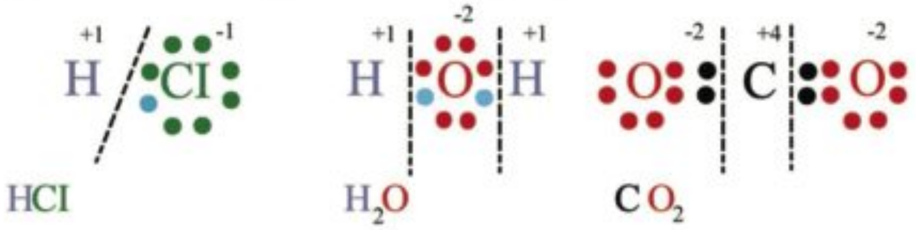


Fig: 2-6

It must be noticed that oxygen in most of its compounds has an oxidation number of (-2) except for few compounds such as peroxides (e.g.)

a) hydrogen peroxide (oxygen water) H_2O_2 in which oxygen has an oxidation number of (-1) Fig (2-7) and in Na_2O_2 (O=-1).

b) Potassium super oxide (KO_2) where oxidation number of O = $-\frac{1}{2}$

c) Oxygen difluoride (OF_2) where oxidation number of O = +2

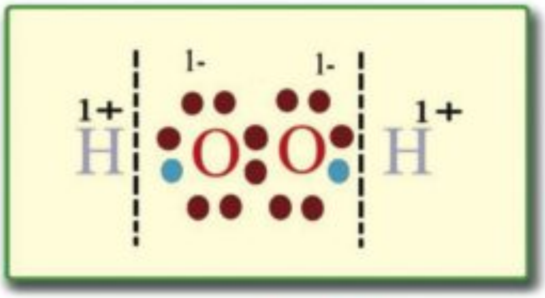


Fig: 2-7

Hydrogen normally has an oxidation of +1 in its compounds except in binary compounds with active metals (metal hydrides), eg. sodium hydride Na H and calcium hydride CaH_2 . Hydrides are ionic compounds in which hydrogen is the negative ion. i.e. if sodium hydride is melted and electrolyzed Hydrogen gas will evolve at the anode the oxidation number of Hydrogen in hydrides (-1) Fig (2-8).

Fig: 2-8

In addition to these basic rules, the below also prove useful:

1. The algebraic sum of the oxidation number of all atoms in a neutral compound is zero.
2. The oxidation number is counted to one atom or ion only in the molecule.
3. The oxidation number of group IA elements is always (+1), group IIA elements is always (+2) and that for group IIIA elements is (+3).

Accordingly, on calculating the oxidation numbers of the atoms in a given compound we start by assigning the oxidation number of the elements of these groups, then we complete the other atoms.

Compound	Na_2O Sodium oxide	Na_2O_2 sodium peroxide	KO_2 Potassium super oxide	CaH_2 Calcium hydride	AlH_3 Aluminum hydride
Total of charges	+2 -2	+2 -2	+1 -1	+2 -2	+3 -3
The atom	Na_2 O	Na_2 O_2	K O_2	Ca H_2	Aℓ H_3
Oxdation no. of each atom	+1 -2	+1 -1	+1 -1/2	+2 -1	+3 -1

4. The algebraic sum of the oxidation number of all atoms in a polyatomic ion is equal to the charge of the ion.

Some common polyatomic ions are:

$(NO_3)^{-1}$ nitrate , $(CO_3)^{-2}$ carbonate , $(SO_4)^{-2}$ sulphate , $(NH_4)^{+1}$ ammonium ion.

The advantage of using oxidation number is that they can help us to tell the type of chemical change occurring to an element during the chemical reaction. For example, in oxidation and reduction reactions, oxidation is the process of losing electrons resulting in an increase of the positive charge; reduction is defined as the process of gaining electrons resulting in a decrease of the positive charge. By following the oxidation number in a chemical reaction, we can recognize the oxidation or reduction process.

Exercise: potassium dichromate reacts with iron (II) chloride (ferrous chloride) according to the equation

$$K_2Cr_2O_7 + 6FeC\ell_2 + 14HC\ell \longrightarrow 2KC\ell + CrC\ell_3 + 6FeC\ell_3 + 7H_2O$$

Explain the type of change (oxidation or reduction) that occurred to chromium and iron in this reaction.

Solution:

To work out the change in oxidation number which occurred, we proceed as follows: To find the oxidation number for chromium in dichromat, we begin by writing down the oxidation number of the elements which are unlikely to be variable. Oxygen usually shows oxidation number(-2) and potassium is always (+1).

Chromium is left at last, because it is a transition element and these element are characterized by having variable oxidation number.

k_2	Cr_2	O_7
(2)x(+l) =+2	2(x) =+2x	(7)x(-2) =-14

From rule 1, the sum of the oxidation numbers should be zero, therefore the value of chromium atoms must be +12 . So the oxidation number of

$$k^{1+}{}_2\,Cr^{6+}{}_2\,O^{2-}{}_7 \longrightarrow Cr^{3+}\,C\ell^{1-}{}_3$$

$$Cr^{6+} \longrightarrow Cr^{3+}$$

one chromium atom is $\frac{+12}{2} = +6$. Similarly, the oxidation number of chromium in the product, $CrC\ell_3$, is (+3). The change in oxidation number can be represented

$Cr^{6+} \longrightarrow Cr^{3+}$

It is clear that the oxidation number of chromium decreases from +6 to +3 i.e it is reduced.

In case of iron: $Fe^{2+}\,C\ell^{1-}{}_2 \longrightarrow Fe^{3+}\,C\ell^{1-}{}_3$

The oxidation number of iron increases from +2 and +3 i.e. the iron has been oxidized.

EVALUATION

QUESTION 1:

Choose the correct answer :

1- Period six in the periodic table contains .... types of elements

a. six b. three c. four d. five

2- Nonmetals are characterized by ....

a. large ionization energy

b. electropositive elements

c. small electron affinity

d. large atomic radius

3- Electronegativity increases across the periods by ....

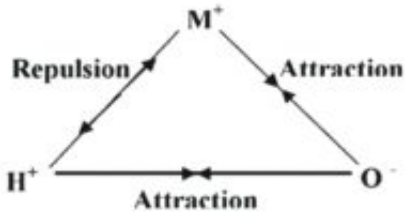
a. increasing atomic radius

b. decreasing the atomic number

c. decreasing atomic radius

d. (a and b) correct.

4- In the shown diagram, if the attraction between O^- and M^+ is greater than that between O^- and H^+, the substance is ionized as a ....



a. base b. acid c. acid and base d. unionized

5- In the previous diagram ,If M^+ is sodium

a- o^- attracts more to hydrogen ion.

b- o^- attracts more to sodium ion.

c- the substance ionized as an acid

d- bound stronger between O^- and sodium

6- One of the following reactions does not represent an oxidation reduction reaction , it is no ....

a. $2P + 5HC\ell O + 3H_2O \longrightarrow 2H_3PO_4 + 5HC\ell$

b. $Zn + 2HC\ell \longrightarrow ZnC\ell_2 + H_2$

c. $Mg + CuSO_4 \longrightarrow MgSO_4 + Cu$

d. $NaOH + HNO_3 \longrightarrow NaNO_3 + H_2O$

7- One of the following reaction represents an oxidation reduction reaction , it is no ....

a. $CuO + H_2SO_4 \longrightarrow CuSO_4 + H_2O$

B. $CaCO_3 + 2HC\ell \longrightarrow CaC\ell_2 + H_2O + CO_2$

c. $Cr_2O^{2-}_7 + 3H_2S + 8H^+ \longrightarrow 2Cr^{3+} + 3S + 7H_2O$

d. $NaC\ell + AgNO_3 \longrightarrow Ag C\ell + NaNO_3$

QUESTION 2:

Choose from column (B) the electronic configuration of the element in column (A) then define the type of element in column (c).

Element	Electronic figuration	Type of element
1-Radon Rn_{86}	a. $7s^1$,	I. inner transition (actinide)
2- Cesium Cs_{55}	b. $6s^2, 5d^6$	II. from 2nd transition series
3- Bromine Br_{35}	c. $6s^2, 5d^{10}, 6p^6$	III. noble element
4- Vanadium V_{23}	d. $4s^2, 3d^3$	IV. 3^{rd} transition series
5-Molybdenum Mo_{42}	e. $6s^2$, $5d^1$, $4f^7$	V. inner transition (lanthanide)
6-Osmium Os_{76}	f. $4s^2, 3d^{10}, 4p^5$	VI. representative from s- block
7-GadoliumGd$_{64}$	g. $5s^2, 4d^4$	VII. first transition series
	i. $6s^1$	VIII. representative from p-block

QUESTION 3:
What is meant by each of the following:

1- Atomic number
2- Reduction
3- Representative element
4- Nobel element
5- Transition element
6- Inner transition element
7- Atomic radius
8- ionization potential
9- Electron affinity
10- Electronegativity
11- Metals
12- Non metals
13- Metalloids
14- Acidic oxide
15- Basic oxide
16- Amphoteric oxide
17- Oxidation number
18- Oxidation

QUESTION 4:
Describe the graduation of the following properties in the periodic table.

a. atomic radius
b. ionization energy
c. electron affinity
d. electronegativity
e. metallic and nonmetallic property
f. acidic and basic property of the elements of the third period

QUESTION 5:
If you know that:

The bond length in hydrogen Fluoride molecules equal (0.94 A°)

The bond length of hydrogen molecule is (0.6 A°), calculate the bond length in fluorine molecule.

QUESTION 6: Explain the differences between:

a. the first and the second ionization energy
b. electron affinity and electronegativity
c. metals and non metals
d. acidic, basic and amphoteric oxide

QUESTION 7:

What is the scientific term for each of the following :

a. half the distance between the centers of two similar atoms in a diatomic molecule.

b. the amount of energy needed to remove the most loosely bound electron completely from an isolated gaseous atom.

c. the amount of energy released when an extra electron is added to neutral gaseous atom.

d. the tendency of an atom to attract the electrons of the chemical bond to itself.

e. the group of elements whose valence shell has less than half its capacity of electrons.

f. the group of elements whose valence shell has more than half its capacity of electrons.

g. A number that refers to the electric charge that the atom would have in the compound.

h. the process of losing electrons associated with increasing the positive charge

i. the process of gaining electrons and associated with a decrease in the positive charge.

QUESTION 8:

Calculate the oxidation number of the mentioned element in following compounds :

a. oxygen in : O_2 - O_3 - Li_2O - Na_2O_2 - KO_2 - OF_2.

b. chlorine in : $NaC\ell$ - $NaC\ell O_4$ - $NaC\ell O_3$ - $NaC\ell O_2$ - $NaC\ell O$.

c. nitrogen in :NH_3 - N_2 - N_2O - NO - HNO_2 - NO_2 - HNO_3.

d. sulphur in : $Na_2S_2O_3$ - K_2S - SO_2 - $NaHSO_3$ - H_2SO_4.

e. manganese in : MnO_2 - $KMnO_4$ - $MnC\ell_2$ - $NaMnO_4$.

QUESTION 9:

By examining changes in oxidation number in the following reactions, decide whether an oxidation or reduction process has taken place.

a) $CO \longrightarrow CO_2$

b) $Cr_2O_7^{2-} \longrightarrow Cr_2O_3$

c) $O_2 \longrightarrow O_3$

d) $NO_2 \longrightarrow N_2O_4$

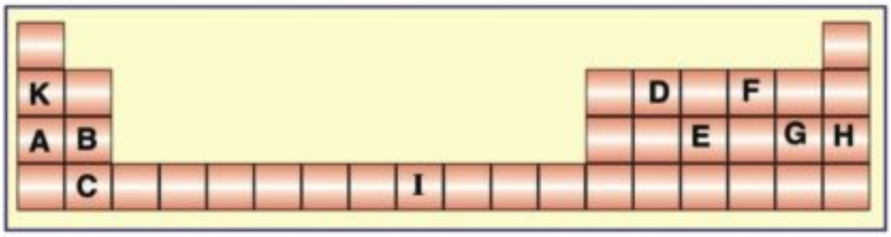
e) $MnO_{4-} \longrightarrow MnO_2$

f) $C\ell O^- \longrightarrow ClO^{3-}$

g) $FeCl_3 \longrightarrow FeCl_2$

QUESTION 10:

The following diagram represents the first four periods of the long form periodic table :



a) Arrange the elements B ,A,G,E in order of decreasing atomic radius.

b) Arrange the elements F,K,D in order of increasing ionization energy.

c) To which block does each of the elements C,H,G,I,D and K belong.

QUESTION 11:

Show by balanced symbolic equation that:

a. Sodium oxide is from basic oxide.

b. Sulphur trioxide is from acidic oxide.

c. Zinc oxide is from amphoteric oxide.

Chapter Three

BONDS AND FORMS OF MOLECULES

Objectives

At the end of this chapter, the student should be able to:

- Construes why atoms tend to form chemical bonds.
- Describe ionic and covalent bonds.
- Define shapes of molecules on the view of valence shell electron pairs repulsion theory.
- Define the type of bond on the basis of electronegativity.
- recognize the electronic theory of valency.
- explain the inadequacies of the octate rule.
- Recognize the formation covalent bond in hydrogen and hydrogen fluoride molecules on the basis of the valence bond theory.
- Recognize the concept of hybridization.
- Explain the overlap of hydrogen and carbon orbitals to form methane molecules.
- Recognize the molecular orbital theory.
- Compare between sigma and pi bonds.
- Compare the type of hybridization of carbon atom in methane, ethylene and acetylene.
- Define the donor and the acceptor atoms in the coordinate bond.
- Recognize the hydrogen bond.
- Draw a diagram to explain the hydrogen bonds formed between water and hydrogen fluoride molecules.
- Conclude the melting point and the hardness of metals according to their valency electrons.

CHEMICAL COMBINATION

We have previously studied the subatomic structure in detail. We have mentioned that the most stable atoms are those of noble gases such as helium, neon, argon ... etc. The atoms of these elements do not undergo any chemical reactions (under normal conditions) with other elements or with each other. Accordingly, they form monoatomic molecules . Referring to the electron structure of these elements, we observe that the outermost energy level is complete with electrons. Consequently , we can conclude that, for an element to be stable, its outermost energy level should also be completed with electrons as Shown in table (3-1)

noble gas	electronic structure
$_{2}He$	$1s^2$
$_{10}Ne$	$[He]2s^2,2p^6$
$_{18}Ar$	$[Ne]3s^2,3p^6$
$_{36}Kr$	$[Ar]4s^2,3d^{10},4p^6$
$_{54}Xe$	$[Kr]5s^2,4d^{10},5p^6$
$_{86}Rn$	$[Xe]6s^2,4f^{14},5d^{10},6p^6$

Table: 3-1

All other elements, except noble gases under normal conditions are reactive to some extent. They undergo chemical reactions to complete their outermost shell by accepting, losing or sharing a number of electrons to acquire an identical electron configuration as that of the nearest noble gas.

As the result of this change in the number of electrons in outermost shells of atoms, bonds are formed or bonds are broken between atoms to form bonds. This is we call a chemical reaction. If chemical bond is not formed or broken among the atoms, there is no chemical reaction.

For example, if iron filing is mixed with sulphure powder, the product.

is not a chemical compound ,this is a mixture, but if this mixture is heated to a high enough temperature to form chemical bonds between them ,we say that a chemical reaction has occurred between iron and sulphur producing the compound iron sulphide.

Valence's electrons play an important role in the nature of the bonds.

Scientist Liwes used a Simple methode by expressing valence electrons with dots (dot diagram) as shown in the following Table:

Group	1A	2A	3A	4A	5A	6A	7A	0
Periode 3	$_{11}Na$	$_{12}Mg$	$_{13}Al$	$_{14}Si$	$_{15}P$	$_{16}S$	$_{17}Cl$	$_{18}Ar$
Electronic configuration	$[Ne]3s^1$	$[Ne]3s^2$	$[Ne]3s^2, 3p^1$	$[Ne]3s^2, 3p^2$	$[Ne]3s^2, 3p^3$	$[Ne]3s^2, 3p^4$	$[Ne]3s^2, 3p^5$	$[Ne]3s^2, 3p^6$
Liwes dod diagram	•Na	•Mg•	•Aℓ• (3 dots)	•Si• (4 dots)	•P• (5 dots)	•S• (6 dots)	:Cℓ• (7 dots)	:Ar: (8 dots)

Liwes named the electron pair found in one of the outer orbitals, which don't share in bond formation by expression [lone pair], and the electron pair responsible for the bond formation by expression the [bond pair] taking into account our new knowledge of the structure of the atom.

We will study two kinds of bonds which are:

Firstly: Chemical bonds **Secondly:** physical bonds

Firstly: Chemical bonds:

I. Ionic bond:

This bond is usually formed between metals and nonmetals. It is known that atoms of metals are characterized by their large volumes. Accordingly, their ionization energies are low. This facilitates the loss of their few electrons of the outermost shell. Metal atoms are then changed to cations with an identical electron structure to the nearest noble gas in the periodic table.

On the other hand, nonmetal atoms are characterized by their small volumes. Accordingly, their electrons affinities are high and this facilitates the gaining of electrons (those lost by metal atoms). Nonmetal atoms are changed to anions with an identical electron structure to the nearest noble gas in the periodic table. consequently, an electrostatic attraction occurs between the positive cations and the negative anions.

It is called the ionic bond. This means that the ionic bond has no materialistic existence and is usually formed between the elements of the two extremes of the periodic table. I.e. the left hand side containing metals and the right hand side containing nonmetals.

Does the difference in electronegativity between the bonded atoms play a role in characteristics of the ionic bond?

To answer this question, examine table (3-2) which compares the bonding of chlorine (which belongs to group seven) to sodium, magnesium and aluminum (metals) which belong to the first, second and third group respectively. All four elements belong to the third period.

Where the electronegativity of chlorine = 3.0

Group	I	II	III
Element	Sodium	Magnesium	Aluminum
Electronegativity	0.9	1.2	1.5
Chloride compound	$NaC\ell$	$MgC\ell_2$	$AIC\ell_3$
Difference in electronegativity	3 - 0.9 = 2.1	3- .1.2=1.8	3 - 1.5 = 1.5
Properties :			
m p/C	810°c	714°c	sublimes
b.p /°C	1465°c	1412°c	sublimes
Electrical conductivity For molten chloride	very good conductor	good conductor	does not conduct

Table: 3-2

It is clear from the table that the differce in electronegativity increases between the bounded elements (horizontally), the ionic character increases. It is proved experimentally that when the difference in electronegativity is more than 1.7, then the formed compound is ionic.

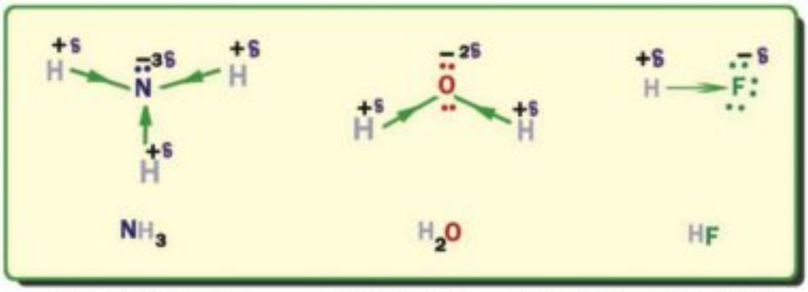
For example, a compound like sodium chloride is definitely ionic in character as reflected in its physical properties : boiling and melting points and the good electrical conductivity when compared to a compound like aluminum chloride in which covalent bond character is more apparent than ionic bond character from an examination of its physical properties.

II. Covalent bond:

We have just mentioned that the ionic bond is formed between atoms with a large difference in electronegativity (more than 1,7). However, when the atoms are of the same electronegativity (atoms of the same elements) or of close electronegativity (till 0.4) (e.g. carbon, e.n = 2.5 and hydrogen, e.n = 2.1), the binding between atoms is achieved by sharing of electron pairs .

This is called a . It is described as a , when the two bonded atoms have equal electronegativity , as (in the bond formed between two fluorine atoms in the diatomic fluorine molecule, or that formed between the hydrogen in the hydrogen molecule). This is because each atom in the molecule has the same ability to attract the two shared electrons of the bond. Thus, the electron pair spends the same time in the vicinity of each atom and the net charge on each atom is zero. However, if the difference in electronegativity is some what high (more than 0.4 but less than 1.7) then the bond is called a . An example of this is the hydrogen chloride bond formed between hydrogen and chlorine in hydrogen chloride (where the electronegativity of chlorine = 3 and for hydrogen = 2.1). The more electronegative chlorine atom has a greater attraction for the electron pair of the covalent bond, i.e. the electrons spend more time in the vicinity of the chlorine atom. As a result, the chlorine atom acquires a partial negative charge and not a complete one (as in the case of chloride ion Cl^-). Because of this unequal sharing of the electron pair towards chlorine, the hydrogen atom acquires a partial positive charge. The hydrogen chloride molecule is then described as a polar molecule. Two other well-known polar molecules are water and ammonia molecules. (See fig.), (3-1).

CO_2 molecule is considered as a nonpolar molecule. inspite of the presence of two polar bonds in it. this is due to the linear shape of its molecule, leads to canceling polar effect of bond by the other [sum of polar pair moment = zero].



Ammonia molecules — Water molecules — Hydrogen fluoride molecules

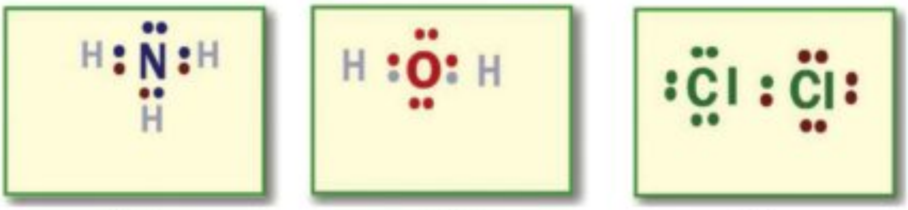
Fig: 3-1

There are more than one theory to explain the formation of the covalent bonds. according to the changing about the concept of electron properties, we shall study some theories to explain formation the covalent bonds.

Theories explain the formation of covalent bond:

a) The Octet Rule or Electronic Theory of Valency

This theory was proposed by Lewis and Kosel in 1916 and it States that: "with the exception of hydrogen, lithium, and beryllium, the atoms of all element tend to reach the octet structure". The covalent bond is formed when a number of electrons from the outermost shell of the two atoms bonded are shared between nuclei. Consequently, the outermost shell of each atom contains eight electrons through the sharing of electron pairs. The outermost shell electrons are represented by dots (•) or crosses (X). Water, chlorine and ammonia molecules can be represented as following:-

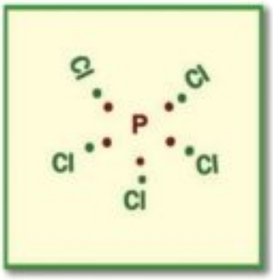


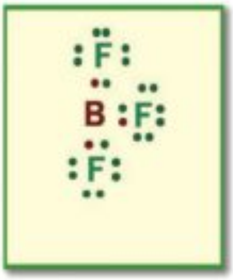
NH_3 Ammonia molecules — H_2O Water molecules — Cl_2 Chlorine molecules

Fig: 3-2

The inadequacies of octet rule:

1- The bonding in many molecules cannot be explained on the basis of the octet rule. In a molecule such as phosphorus pentachloride, the phosphorus atom is surrounded by ten electrons and not eight as the theory proposed. Also in the boron trifluoride molecule the boron atom is surrounded by only six electrons.





PCl_5 BF_3

Fig: 3-3

2- The simple representation of the covalent bond as being just a shared pair of electrons, is not sufficient to explain many of the properties of molecules such as the stereostructure and the angles between bonds.

b) The valence Bond Theory (V.B.T.)

This theory was based on the conclusions of quantum mechanics that considers the electron not only as negative particle that moves in a definite orbit, but as a material particle with wave properly which can exist in any position in the space surrounding the nucleus. The V.B.T explain the formation of the covalent bond as a result of the overlap of an atomic orbital of one atom which contains a single electron with a similar orbital of another atom. The hydrogen molecule is formed as a result overlap of the electron of the Is orbital of each atom as shown in fig. (3-4)

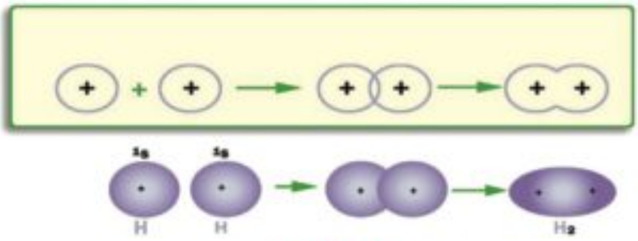


Fig: 3- 4

For another molecule such as hydrogen fluoride, one of the 2p orbitals of fluorine atom has a single electron overlapped with the electron of 1s orbital in the hydrogen atom (See fig) .(3-5)

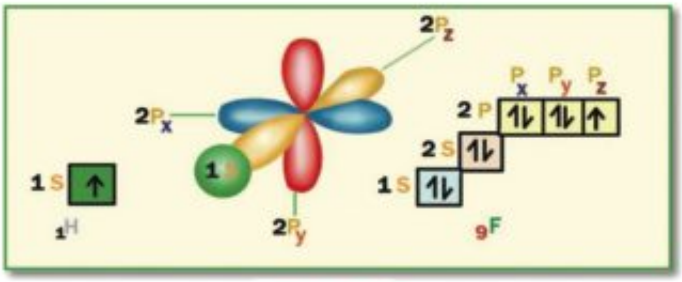


Fig: 3-5

How does the valence bond theory explain the structure of methane molecule?

On studying the bonds in methane molecule CH_4 ,the carbon atom is binded with four hydrogen atom by four symmetrical bonds in length and strength ,and the molecule take the tetrahedron shape and the angles between the bonds = 109.5°

The valence bond theory explain these results as the following:

In the ground state, the carbon atom has only two orbitals with two single electrons which are capable of forming covalent bonds. However, carbon forms four covalent bonds (as in methane CH_4) and not just the two bonds predicted by pairing the 2p electrons . Therefore, the carbon atom must have four single electrons, which can be paired to form compounds achieved by exciting one electron from the 2s orbital electron pair to the vacant 2p orbital. Then the atom is said to be in an excited state. This excitation takes place when the atom acquires a given amount of energy. Now, the carbon atom has four single electrons, but are these four single electrons equivalent as the four bonds in methane are equivalent? The answer must be no since one electron is in a lower 2s orbital.

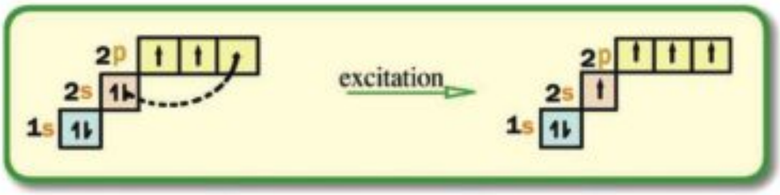


Fig: 3-6

Hybridization:

The hybridization process is the overlap between two different orbitals or more of the same atom to produce a number of new orbitals called hybridized orbitals.
It is obvious that in carbon atom the electron of the 2s orbital differ from the three electrons of the three 2p orbitals . (i, .e differ from them in energy and in the form of the orbital) .

Therefore, there must be a further chance that occurs to the atomic orbital in the carbon atom resulting in the formation of four equivalent orbitals. This process is termed hybridization . The hybridization process occurs between the orbital of the same atom which are close to each other in energy e.g.2s and 2p or 3d, 4s and 4p to produce a number of hybridized orbitals that have become a part of hybridization process.

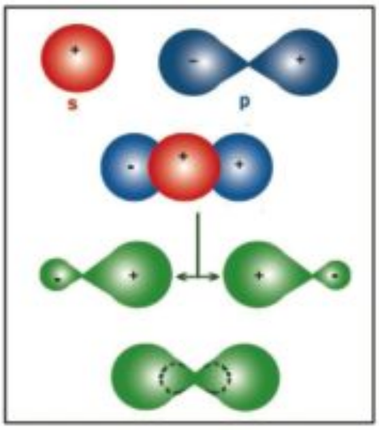


Fig: 3-7

The shape of the hybridized orbitals differ from these of the pure atomic orbital forming them. The hybridized orbitals must protrude to the outside to be more capable of overlapping than the pure atomic orbital (See fig 3-7).

The experimental data has shown that in methane molecule the carbon atom is bound to four hydrogen atoms through four equivalent bonds equivalent in both strength and length. The molecule takes the form of tetrahedron in which the angles between the four bonds are 109.5°.

The theory explains these angles as the following:

Each of the hybridized orbitals in the carbon atom contains a negative electron.

These orbitals must go apart as far as possible from the other orbital to decrease the repulsive forces between orbital. When the angles between orbitals are 109° 5\`

they will be more stable (less repulsive) compared to angles of 90° (an alternate structure), to form the methane molecule.

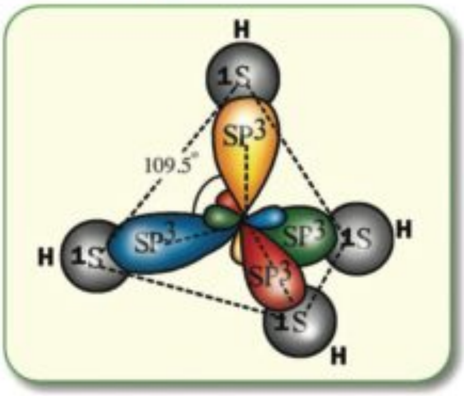


Fig: 3-8

The four equivalent electrons of the four hybridized orbitals of the carbon atom can overlap with the four 1s electrons of the four hydrogen atoms . (fig 3-8).

The following aspects must be considered in the hybridization process:

1. Hybridization occurs between the orbitals of the atom itself.
2. It occurs between orbitals of close energy .
3. The number of hybrid orbitals equals the number of pure orbitals undergoing hybridization with their symboles.

Another example using sp^2 hybridization to explain the hybridization of the carbon atom in the ethylene (ethene) molecule C_2H_4 . If each carbon atom in the molecule is sp^2 hybridized, then the molecule can be formed as shown in fig.(3-11)

It is clear from fig. that after the two carbon atoms have been excited, the three orbitals of 2s, $2p_x$ and $2p_y$ are hybridized forming three hybridized orbital each of which is named an sp^2 hybrid orbital.

To overcome the repulsive forces between the orbitals .They are directed in space as far apart as possible thus the angles between them are 120°.

It is seen that the $2p_z$ orbital of each carbon atom is not involved in the hybridization process. These two orbitals lie perpendicular on the plane of the three sp^2 orbitals of each carbon atom. Two types of overlap between carbon atoms can now occur. **These are:**

a- Two sp^2 hybrid orbitals on each carbon atom overlap with the 1s orbital of two hydrogen atoms, forming the two C-H bonds.

The third sp^2 orbital

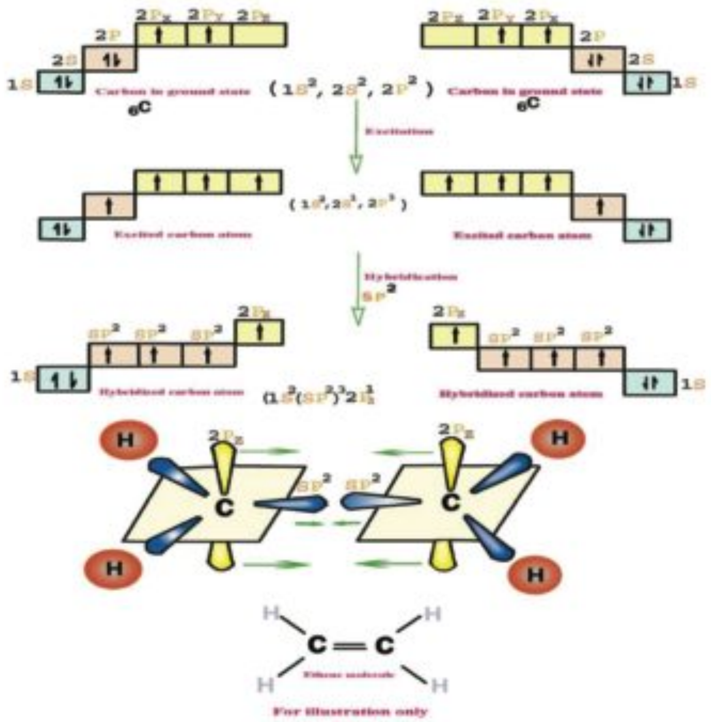


Fig: 3-9

on one carbon atom overlap with the similar one on the other atom forming the C-C bond .This type of bond is known as sigma bond in molecular orbitals theory.

b- The two $2p_z$ orbitals on the two carbon atoms overlap to form another type of bond between them, known as (pi) bond.

The sigma bond (σ)

This is formed as a result of the overlap of atomic orbitals along an axis, i.e. the overlapped orbitals are on one line or "collinear overlap". An example of sigma bond formation is the overlap of the hybrid sp^2 orbitals of carbon atom with the 1s orbitals of hydrogen atom, to form C-H bond or the overlap of a carbon sp^2 hybrid orbital with the sp^2 in molecule orbital of the other carbon atom to form a C-C bond as in , C_2H_4 ethylene molecule as in fig. (3- 9).

The Pi bond (π)

This is formed as a result of the overlap of atomic orbitals side by side, i.e. the overlap is between two parallel orbitals or "collateral overlap". An example of a Pi bond is the overlap of the $2p_z$ atomic orbital of one carbon atom side by side with its similar orbital in the other carbon atom as in the C_2H_4 molecule as in fig. (3-9).

In the acetylene molecule, C_2H_2 the two carbon atoms are linked together by one sigma bond and two pi bonds as shown in fig. (3-10).

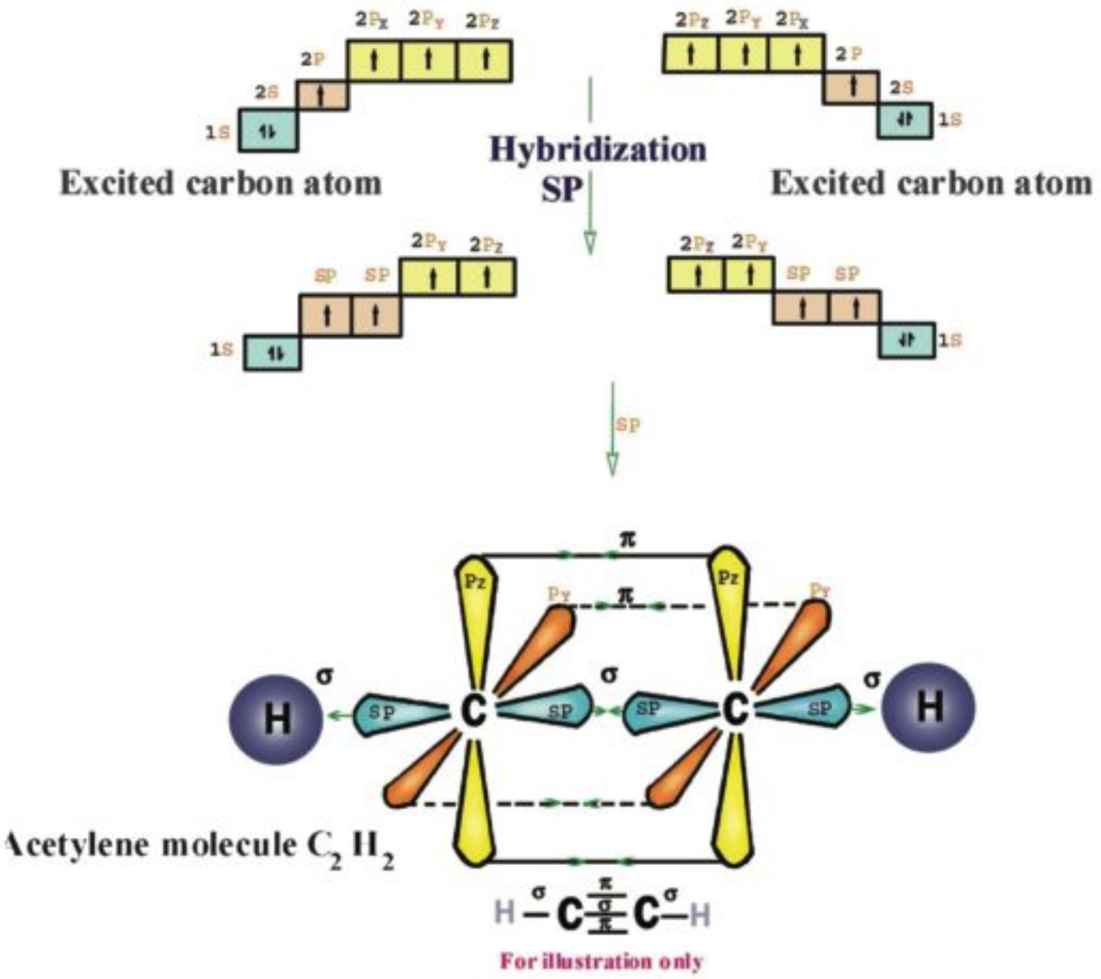


Fig: 3-10

It is clear from fig. (3-10) that after the excitation process has occurred in the two carbon atoms, hybridization occurs between one orbital ($2p_x$) from sublevel (p) orbital and the (2s) orbital producing two sp hybridized orbitals.

I(2s) +1 (2p) ⟶ 2(sp)

To overcome the repulsion forces between the two hybridized orbitals , they are directed in space as far apart as possible, forming an angle of 180°.

We observe that in each carbon atom there are two atomic orbital remaining, i.e. $2p_z$ and $2p_y$. they arc not involved between the orbital of each carbon atom then the overlap between the hybridized orbitals of each carbon atom forms the following bonds:

a) One sigma bond produced from the "head to head" overlap between one sp orbital on each carbon atom forming the C-C bond. Sigma bonds are also formed by the remaining orbitals on each carbon atom with the 1s atomic orbital on each hydrogen atom.

b) Two pi bonds between carbon atoms are produced by the "side by side" overlap of the two parallel $2p_y$. orbitals and the other two parallel $2p_z$ orbitals ,of two carbon .

Shapes of molecules according to valence shell electron pairs repulsion theory (VSEPR).

Shapes of molecules differ according to (Free and bonded) electron pairs which are found in the orbitals of central atom in the covalent molecule or distributed in the space around the central atom of the molecule. where the repulsion between them is the minimum. (very small), to form the most stable shape of the molecule as (VSEPR) theory said.

The following table shows the shape of some molecules according to the (VSEPR theory).

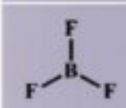
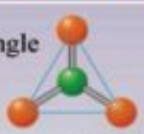
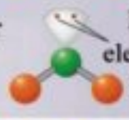
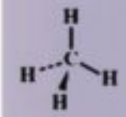
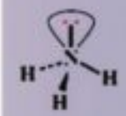
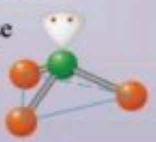
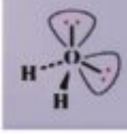
Examples of pacticles		Shape of the solid part		Arrangement of the electron pairs	Electron pairs		
					Free	Bonded	Resultant
F—Be—F	BeF_2	Linear Ax_2		Linear	0	2	2
	BF_3	Plane triangle Ax_3		Planner Triangle	0	3	3
	SO_2	Angular Ax_2E	 Free electrons pair		1	2	
	CH_4	Tetrahedron Ax_4		Tetrahedral	0	4	4
	NH_3	three - base pyramid Ax_3E			1	3	
	H_2O	angular Ax_2E_2			2	2	

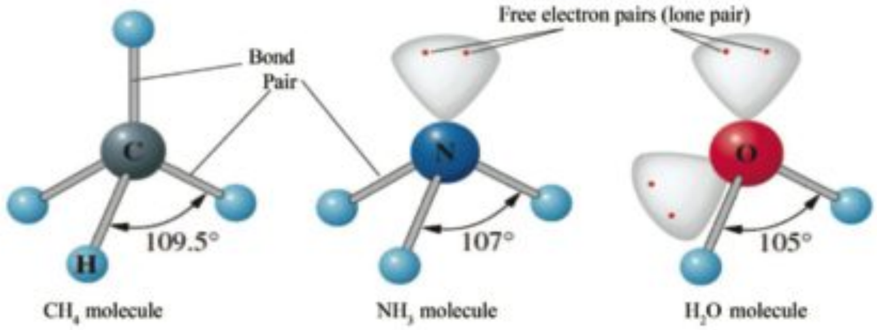
Fig: 3-11

● X = bonded atom to the central atom ● A = central atom

The lone pair of electrons control the value of bond angles in the molecule, because the lone pair of electrons is bound to nucleus of centeral atom from one side and spread to the space from other side. While the bond pair is bound from both sides with the nuclei of the two bonded atoms. valence shell electron pair repulision theory VSEPR Increasing the number of lone pair of electrons in central atom of the molecule leads to increasing the repulsive force between them. which cause decrease value of angles between covalent bond of molecule and in general, the repulsion is between:

(Lone pair , lone pair) > (Lone pair, bond pair) > (bond pair, bond pair)

How can you explain decreasing the values of covalent bond angles in water than ammonia than methane?



Free electron pairs control in determining the values of angles among the bonds in covalent molecules.

Fig: 3-12

The lone pair of electrons control the values of bond angles in covalent molecules.

III. The molecular orbital theory (MOT)

The valence bond theory considered the formation of covalent bonds in the molecule as a result of the overlap of some atomic orbitals in the combined atoms. The rest of the atomic orbitals which did not take part in the formation of bonds remain as they were in the free atoms. However, the molecular orbitals theory considered the molecule as one unit (or a big atom with multinuclei) in which all the atomic orbitals of the combined atoms are mixed or hybridized forming molecular orbitals. As the atomic orbitals have the symbols s, p, d .., etc so the molecular orbitals have symbols: sigma (σ), pi (π) and delta (δ)... etc

IV. Coordination bond

The coordinate bond is a type of covalent bond which differs only in the origin of the electron pair which is shared between the atoms involved in the covalent bond. The electron pair forming the ordinary covalent bond is produced by both of the two combined atoms contributing one electron to the bond. Whereas, the electron pair of the coordinate bond is a lone electron pair occupying one orbital found in one atom called the donor atom. This lone pair is donated to another atom having a vacant orbital i-e,

needing this electron pair to acquire the stable electron configuration. This atom is called the acceptor atom. An example of a coordinate bond is that

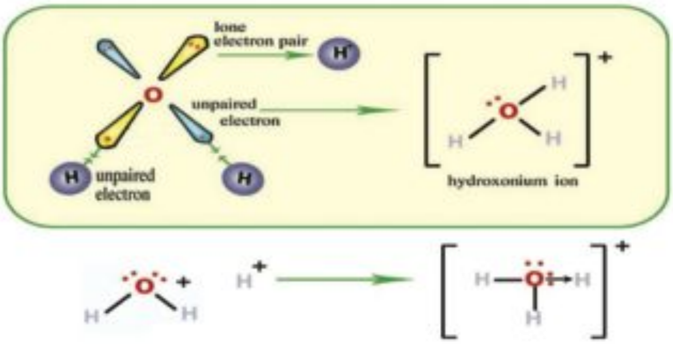


Fig: 3-13

formed between the neutral water molecule and the positive hydrogen ion produced on dissolving acids in water, thus forming the hydroxonium ion (H_3^+0) . See fig: (3-13)

The coordinate bond is also formed in the ammonium ion (NH_4^+) when the proton H^+ accepts the lone electron pair from the nitrogen atom of the ammonia molecule.

Secondaly: physical bonds:

I. Hydrogen bond

It is known that water boils at 100° C. This temperature is considered to be high for a compound of low relative molecular mass (18 g/mole). If we compare this with the boiling point of hydrogen sulphide (of molecular weight 34 g/mole) we find that it is -61° C. anomalous of boiling point of water is due to higher electronegativity of oxygen (3.5) than hydrogen (2.1), as a result of the difference in electronegativity. So the water molecule is said to be a polar molecules . The oxygen atom in water carries a partial negative charge, while the two hydrogen atoms carry a partial positive charge. As a result of the separation of charges in the water molecules, They are attracted together by what is known as the hydrogen bond, or the hydrogen bridge .

Hydrogen bond:

Bond formed between hydrogen atom binds by polar bond [like N-H, O-H, H-F] with high electronegative bonded atom like [N,O, F].

The hydrogen atom becomes a bridge between two oxygen atoms of high electronegativity .The molecules get near to each other, so that we can say that the hydrogen atom binds the water molecules together see fig (3-15)

Hydrogen bonds are a type of intermolecular force of attraction.

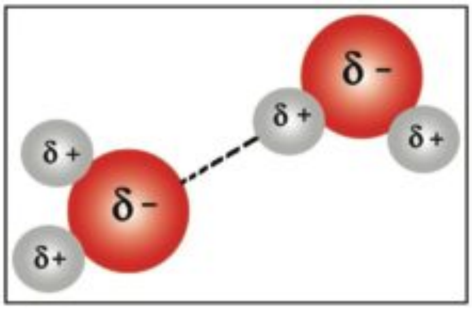


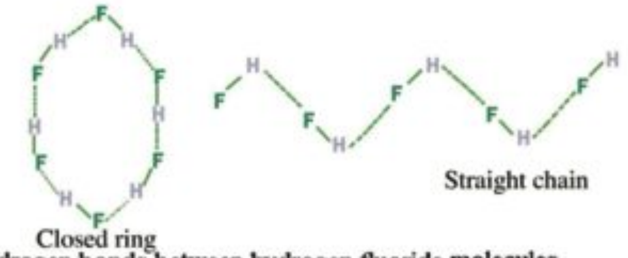
Fig: 3-14

Now we can account the boiling point of water by the fact that the heat energy is used up in dissociating the hydrogen bonds between water molecules .

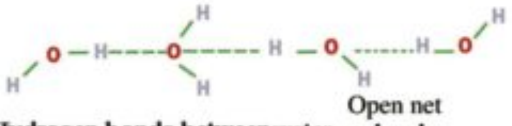
Although the hydrogen bond clearly affects physical properties its bond strength is far weaker than the normal chemical bonds. The following table shows the difference between the covalent and hydrogen bonds.

	Bond length	Bond strength
Covalent bond	1 A°	418 kj/mol
Hydrogen bond	3 A°	21 kj/mol

It is clear that the hydrogen bond is much weaker than the covalent bond and is longer than it. when hydrogen lies on straight line with the polar covalent bond, as in water molecule. (H_2O) and hydrogen fluoride molecule (H-F) and as the electronegativity difference between hydrogen atom and the other atom bind to it increases polarity increase. For covalent molecules which have hydrogen bonds between them, the hydrogen bonds may be of several forms, i.e. The molecules may form a straight line, a closed ring or an open net of hydrogen bonds as in hydrogen fluoride and water . See fig- 3-15



Hydrogen bonds between hydrogen fluoride molecules



Hydrogen bonds between water molecules

Fig: 3-15

II. Metallic bond

Each metal has a crystal lattice with a definite form in which the positive metal ions take a certain arrangement. The outermost shell electrons of each atom are associated together forming an electron cloud be with free movement to bind this great collection of positive metal ions.

This means that the metallic bond is produced from the electron cloud be of valence electrons which decreases the repulsive forces between the positive metal ions in the crystal lattice.

These free valence electrons account for the good electrical and thermal conductivity of metals. The number of valency electrons in the metal atom plays a role in the strength of the metallic bond. As the number of valence electrons increases in the metal atoms, the atoms become more strongly bound and accordingly the metal becomes more hard and has higher melting point. This is evident when comparing the properties of sodium, magnesium and aluminum (the metals of the third period).

Metal	Number of the outermost	Hardness mohw scale	Melting point °C
Na	1	(soft) 0.5	98 °C
Mg	2	(mild) 2.5	650 °C
A1	3	(hard) 2.75	660 °C

Evaluation

QUESTION 1:-

What is meant by:

1- Chemical reaction

2- Ionic bond

3- Covalent bond

4- Coordinate bond

5- hydrogen bond

QUESTION 2:-

Choose the correct answer:

1- Three elements of atomic numbers C_{11}, B_{10}, A_9: the possible combination takes place between the atoms:

a- B with C   b- A with B

c- b with it self   d- C with A

2- When two atoms of an element of atomic no 9 combine together to form a molecule, the formed bond is:

a- metallic   b- coordinate   c- ionic   d- covalent

3- The bond in hydrogen fluoride molecule is polar covalent, because the atoms are different in :

a- location in periodic table

b- Electron affinity

c- Electronegativity

d- Ionization potential

4- SP hybridized orbitals are characterized by :

a- They are three in number .

b- They are two orbitals .

c- They are linear in shape .

d- (band c) are correct.

5- In acetylene molecule:

a- The two carbon atoms are linked by one sigma and one pi bonds .

b- The two carbon atoms are linked by one sigma and two pi bonds.

c- The two carbon atoms are linked by sp hybridized orbitals.

d- (b and c) are correct.

6- (SP^3) hybridized orbitals produced from the overlap of:

a- one s and two p orbitals.

b- two s and two p orbitals .

c- one s and three p orbitals .

d- one s and one p orbitals

7- When two oxygen atoms combine together to form a molecule , this process takes place by :

a- each atom shares by one electron to form one covalent bond.

b- one atom donates a lone pair of electrons to the other one .

c- each atom shares by two electrons.

d- a polar covalent bond is formed between the two atoms

QUESTION 3:-

Give reason for:

1- However oxygen and sulphur are located in group six in the periodic table ,the boiling points of their hydrogenated compounds have different boiling points , i-e water boils at 100°c whereas hydrogen sulphide boils at -61°c.

2- The negative fluoride ion and the positive sodium ion are isoelectronic.

3- A coordinate bond is formed in ammonium molecule.

4- CO_2 molecule is nonpolar, in spite of the presence of two polar covalent bonds in it.

5- Value of bond angle in ammonia molecule is less than in methane molecule.

QUESTION 4:-

Answer the following:

a- Explain the expected bond in each of the following compounds.

KC1 - NO - SO_2- HC1 - CaO .

b- Arrange the following bonds according to their polarity.

H-Cl , C=O , H-H , N-O , P-Cl .

QUESTION 5:-

Write the scientific term for each of the following:

1- A bond formed by the collateral overlap of two atomic orbitals.

2- A bond formed when hydrogen atom is located between two atoms of high electronegativity.

3- Ion formed when hydrogen ion links with water molecule.

4- Bond formed between two atoms the difference in electronegativity between them is zero.

QUESTION 6:-

Explain using Lwies diagram bonding in:

1- Sodium with chlorine to from formula unit ($NaC\ell$).

2- Nitrogen with hydrogen to form ammonia molecule (NH_3).

QUESTION 7:-

Compere between each two pairs of the following in terms of:

(Sterio structure , Number of lone pair and bond pair of electrons)

A- CH_4 , BeF_2 B- SO_2 - BF_3

QUESTION 8:-

Redraw structure of hydrazine molecule (N_2H_4) in front of you using dot diagram for electron pairs (lone and bonded):

```
        H     H
        ··    ··
H  :    N  :  N  :  H
        ··    ··
```

QUESTION 9:-

Define The sterio Structure for the molecule which contains two atoms bonded to the central atom and 1 lone pair with writing abbreviation expressing it.

QUESTION 10:-

Find number of atoms bonded to the central atom, lone pair and also arrangement of electron pairs in the molecule which has abbreviation AX_2E.

Chapter Four

The representative elements of some regular groups in the periodic table

Objectives

At the end of this chapter, the student should be able to:

- Recognize the alkali metals and their electronic configuration .
- Recognize the general characteristics of alkali metals
- Conclude methods of extraction of alkali metals from their ores .
- Explain the general properties of sodium hydroxide.
- Practice some experiments for the identification of basic radicals.
- Describe the methods of preparation of sodium carbonate in both lab and industry .
- Recognize the general characteristics of the element of group five and their electronic configuration.
- Define the oxidation number of nitrogen in different compounds.
- Recognize methods of preparation of nitrogen gas in the lab and its physical and chemical properties.
- Know how ammonia is prepared in laboratory and industry .
- Carry out an experiment to identify ammonia gas.
- Compare different types of nitrogenous fertilizers .
- Know how to prepare nitric acid in lab.
- Recognize the criteria of nitric acid .
- Distinguish between nitrate and nitrite salts.
- Know the economic importance of the fifth group elements.
- Consider safety rules in laboratory.
- Recognize scientist efforts in serving humanity.

The representative elements of some regular groups

From our study of the periodic table, we have seen that one of the aims of the classification of elements is to facilitate their study in a regular way . We shall study the elements, in some of the main groups, which are known as A-groups. In these main groups there is a great regularity and graduation in their properties, the property which is not found in the transition elements of B-groups.

In this chapter, we shall offer a more detailed study of the chemistry of these elements, and interpret the observed behavior using the theoretical principles that we have studied before(such as atomic radii, ionization energies and the electronegativity). The chemical and physical properties of these elements can be explained using these principles.

Firstly: Elements of s-Block

Example: Elements of group (IA) (Alkali metals)

Elements of this group are known as alkali metals (forming alkalis) Moslem scientists gave the name "Al-Kale" to both sodium and potassium compounds. This name was borrowed by Europeans, becoming " Alkali" Then, this word was used to encompass all elements of this first group.

The alkali metals group comprises six elements

Element	Symbol& At.no	Electronic distribution according to building up principle
Lithium	$_{3}Li$	$[He]_2\ 2S^1$
Sodium	$_{11}Na$	$[Ne]_{10}3S^1$
Potassium	$_{19}K$	$[Ar]_{18}4S^1$
Rubidium	$_{37}Rb$	$[Kr]_{36}5S^1$
Cesium	$_{55}CS$	$[Xe]_{54}\ 6S^1$
Francium	$_{87}Fr$	$[Ra]_{86}7S^1$

Abundance of alkali metal in nature:

Sodium and potassium are abundant elements in the earth crust. They are the 6^{th} and 7^{th} most common elements in the earth's crust. The most important ore of sodium is Rock Salt ($NaC\ell$).

The most important ore of potassium is potassium chloride, which is found in sea water, and also in carnallite deposits: $KC\ell.MgC\ell_2\ 6H_20$

Other metals of this group are rare, e.g. francium.(which is a radioactive element that was discovered in 1946) as a product of the disintegration of actinium .

$$^{227}_{89}Ac \longrightarrow ^{223}_{87}Fr + ^{4}_{2}He$$

Since the amount of francium formed in this decay process is very small, all we know about this element is its atomic number, and its approximate atomic weight. It is a radioactive element whose half life period is only twenty minutes.

General properties of the first group elements:

1- All elements of this group are characterized by the presence of one single electron on the outer energy level, (ns^1)Accordingly:

a) Each element of this group is at the beginning of a new period in the periodic table.

b) The oxidation number of all group (1 A) elements in their compounds is (+1).

c) Due to the ease of losing valence electron they are chemically very active. So the first ionization potential is less than the ionization potential of any other element in the period but the second ionization potential is very large. Since the second electron will be removed from a complete or saturated energy level which is stable.

d) Most of their compounds are ionic, the ion of each element is identical in electron structure to the noble gas which precedes it.

e) Since the number of electrons in the outer energy level in the metal atom is one of the factors which controls the strength of the metallic bond. Since metals of the first group have only one electron in tlhe outer energy level (valency orbit),these metals are characterized by a small attraction between their atoms ,so that they are the most malleable metals, with the lowest melting and boiling points.

f) They are strong reducing agents.

2) Atoms of this group have the largest known volume, so any metal of this group has the largest volume of any atom in its period. The volume of the atom increases down the group i.e. with increasing atomic number. Due to the increase of the volume of the atom , alkali metals show the following properties :

a) An increase in the radius of the atom , decreases the force of attraction between the valency electron and nucleus. This electron can be lost easily, therefore these metals are considered of the highest electropositive and chemical reactivity.

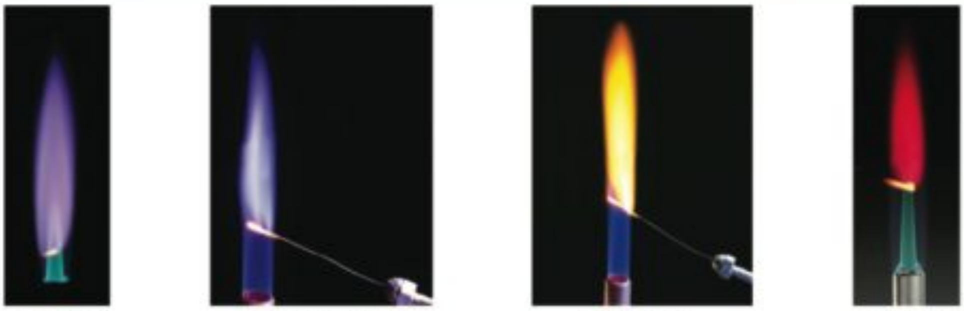
b) The phenomenon of having a large volume of the atom and small ionization energy, is used in photo-electric cells as in **potassium and Caesium**. These elements when exposed to light, lose electrons from the outer surface of the metals; this is known as the photo electric phenomenon.

C) Low densities.

d) These elements have very low electonegativity , compared with other elements when combining with other elements they form strong ionic bonds.

3- When the electrons of these elements are excited to higher energy levels. They give the characteristic colours of these elements .

Element	Colour
Lithium	Crimson
Sodium	Golden yellow
Potassium	Pale violet
Cesium	Bluish violet



(Flam test for metals of 1st group)

fig: 4-1

This property is used in the dry test of these elements (flame test) in their compounds as the following:

The platinum wire is dipped in a concentrated hydrochloric acid to clean it. Then dip the platinum in the unknown salt and expose it to the non- illuminant Bunzen flame. The flame will acquire the specific colour of cation.

4 - Due to the high activity of these elements they are kept out of air and humidity. They are stored under liquid hydrocarbons e.g. kerosene.

5- Action of atmospheric air :

All elements of this group are chemically active. They are oxidized easily in air to lose their metallic luster forming the oxide .Lithium can react with nitrogen of the air giving lithium nitride, which itself reacts with water to evolve ammonia .

$$\underset{(s)}{6Li} + \underset{(g)}{N_2} \xrightarrow{\Delta} \underset{(s)}{2Li_3N}$$

$$\underset{(s)}{Li_3N} + \underset{(\ell)}{3H_20} \longrightarrow \underset{(aq)}{3LiOH} + \underset{(g)}{NH_3}$$

6- With water :

Elements of this group are considered to be very active compared with any other known metal. They are at the top of the electrochemical series. They can replace the hydrogen of water .The reaction is accompanied by the evolution of a large amount of energy,

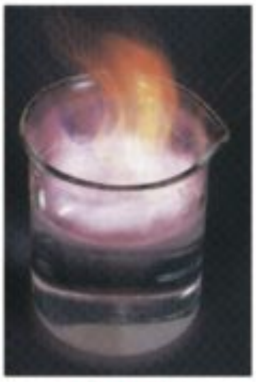


fig: 4-2

reaction of sodium with water

which leads to the burning of the hydrogen evolved. The reaction becomes more vigorous down the group and with caesium an explosion occurs.

$$\underset{(s)}{2Na} + \underset{(\ell)}{2H_2O} \longrightarrow \underset{aq}{2NaOH} + \underset{(g)}{H_2}$$

Can we extinguish burning sodium fires with water?

7- With oxygen : The trend in reactivity in elements of this group is clear when it reacts with oxygen. When these elements are burnt ,they give three types of oxides. Lithium gives the normal oxide(Li_2O)in which the oxidation number of oxygen is (-2). Sodium gives sodium peroxide (Na_2O_2), which gives peroxide ion $(O_2)^{-2}$ the oxidation number of oxygen in this compound is (-1).

Potassium, rubidium, and caesium give super oxides e.g. oxidation number of oxygen is (-½).

$$\underset{(s)}{4Li} + \underset{(g)}{O_2} \xrightarrow{180^\circ C} \underset{(s)}{2Li_2O}_{\text{(Lithium oxide)}}$$

$$\underset{(s)}{2Na} + \underset{(g)}{O_2} \xrightarrow{300^\circ C} \underset{(s)}{Na_2O_2}_{\text{(Sodium peroxide)}}$$

$$\underset{(s)}{K} + \underset{(g)}{O_2} \xrightarrow{300^\circ C} \underset{(s)}{KO_2}_{\text{(potassium superoxide)}}$$

Both peroxides and super oxides compounds act as a strong oxidizing agents, as they react with water and acids giving hydrogen peroxide and oxygen in case of super oxide.

$$\underset{(s)}{Na_2O_2} + \underset{(\ell)}{2H_2O} \xrightarrow{\text{cold}} \underset{(aq)}{2NaOH} + \underset{(aq)}{H_2O_2}$$

$$\underset{(s)}{2KO_2} + \underset{(\ell)}{2H_2O} \xrightarrow{\text{cold}} \underset{(aq)}{2KOH} + \underset{(aq)}{H_2O_2} + \underset{(g)}{O_2}$$

$$\underset{(s)}{Na_2O_2} + \underset{(aq)}{2HC\ell} \longrightarrow \underset{(aq)}{2NaC\ell} + \underset{(aq)}{H_2O_2}$$

$$\underset{(s)}{2KO_2} + \underset{(aq)}{2HC\ell} \longrightarrow \underset{(aq)}{2KC\ell} + \underset{(aq)}{H_2O_2} + \underset{(g)}{O_2}$$

Potassium super oxide reacts with carbon dioxide giving oxygen.

$$\underset{(s)}{4KO_2} + \underset{(g)}{2CO_2} \xrightarrow[\text{catatyst}]{CuC\ell_2} \underset{(s)}{2K_2CO_3} + \underset{(g)}{3O_2}$$

This reaction is used in replacing carbon dioxide by oxygen in closed atmospheres such as submarines or airplanes (which fly at very high altitudes). Exhale gases which contain a large ratio of carbon dioxide are passed though filters containing potassium super oxide and the catalyst. The normal oxide of these elements is X_2O, where X is the symbol of the element can be prepared by dissolving the metal in liquefied ammonia and then adding a calculated amount of oxygen . With the exception of Li_2O, the group 1A oxides react with water to give strongly alkaline solutions

8- With acids :

These metals can replace the hydrogen in acids. It is such a vigorous reaction.

$$\underset{(s)}{2Na} + \underset{(aq)}{2HC\ell} \longrightarrow \underset{(aq)}{2NaC\ell} + \underset{(g)}{H_2}$$

9- With hydrogen : Alkali metals react with hydrogen giving hydrides.

$$\underset{(s)}{2Li} + \underset{(g)}{H_2} \xrightarrow{\Delta} \underset{(s)}{2LiH} \text{ Lithium hydride}$$

$$\underset{(s)}{2Na} + \underset{(g)}{H_2} \xrightarrow{\Delta} \underset{(s)}{2NaH} \text{ Sodium hydride}$$

Hydrides are reducing agent , react with water and hydrogen gas evolve

$$\underset{(s)}{LiH} + \underset{(l)}{H_2O} \longrightarrow \underset{(aq)}{LiOH} + \underset{(g)}{H_2}$$

Hydrides are ionic compounds in which hydrogen has oxidation no.(- 1)

10- With halogens :

Alkali metals react vigorously with halogens forming very stable ionic halides.

$$\underset{(s)}{2Na} + \underset{(g)}{C\ell} \longrightarrow \underset{(s)}{2\ NaC\ell}$$

$$\underset{(s)}{2K} + \underset{(l)}{Br} \longrightarrow \underset{(s)}{2\ KBr}$$

11 - With other non-metals :

Hot alkali metals can react directly with sulphur and phosphorus

$$\underset{(s)}{2Na} + \underset{(s)}{S} \xrightarrow{\Delta} \underset{(s)}{Na_2S} \text{ sodium sulphide}$$

$$\underset{(s)}{3K} + \underset{(s)}{P} \xrightarrow{\Delta} \underset{(s)}{K_3P} \text{ potassium phosphide}$$

12 - The action of heat on alkali metal oxygenated salts:

Alkali metal oxygenated salts are thermally stable

a) All alkali metal carbonates do not decompose when heated ,except lithium carbonate which decomposes at 1000 C .

$$\underset{(s)}{Li_2CO_3} \xrightarrow{1000\ C^o} \underset{(s)}{Li_2O} + \underset{(g)}{CO_2}$$

b) Alkali metal nitrates decompose partially giving a metal nitrite and oxygen.

$$\underset{(s)}{2NaNO_3} \xrightarrow{\Delta} \underset{(s)}{2NaNO_2} + \underset{(g)}{O_2}$$

An explosion takes place when potassium nitrate decomposes by heat. So, it is used in the manufacture of bombs .

'N .B.

Sodium nitrate is not used to make bombs because it is deliquescent i.e. It absorbs water vapour from air .

Extraction of alkali metals from of their ores

Elements of this group are the most powerful reducing agents, i.e. these metals are easily lose their valence electrons .

These elements are not found in nature in a free state, but rather in the form of ionic compounds like sodium chloride ($NaC\ell$). Usually the preparing of these metals involves the electrolysis of its molten (fused) halide in the presence of a flux substance to decrease the melting point of these halides.

Ex

at cathode $2Na^+ + 2e^- \longrightarrow 2\ Na$

at anode $2C\ell^- \longrightarrow C\ell_2 + 2\ e^-$

Commonly used sodium compounds:

(1) Sodium Hydroxide Na OH

A) Properties :

1- A white hygroscopic solid compound (absorb water vapour from atomspheric air).

2- It has a soapy touch and a corrosive effect on skin .

3- It dissolves easily in water forming an alkaline solution through an exothermic dissolution .

4- It reacts with acids forming the sodium salt of the acid and water.

$$\underset{(aq)}{NaOH} + \underset{(aq)}{HC\ell} \longrightarrow \underset{(aq)}{NaC\ell} + \underset{(\ell)}{H_2O}$$

$$\underset{(aq)}{2NaOH} + \underset{(aq)}{H_2SO_4} \longrightarrow \underset{(aq)}{Na_2SO_4} + \underset{(\ell)}{2H_2O}$$

B) Uses :

1- Sodium hydroxide is used in many important industries such as: Soap, synthetic silk and paper.

2- It is used to purify petroleum from the acidic impurities .

3- Detection of basic radicals (cations):-

a- Detection of copper (II) (Cu^{++}).

salt solution + sodium hydroxide solution gives a blue precipitate of copper (II) hydroxide - turns black on heating.

$$\underset{(aq)}{CuSO_4} + \underset{(aq)}{2NaOH} \longrightarrow \underset{(aq)}{Na_2SO_4} + \underset{\text{blue ppt (s)}}{Cu(OH)_2}$$

$$\underset{(s)}{Cu(OH)_2} \xrightarrow{\text{heat}} \underset{(v)}{H_2O} + \underset{\text{black ppt (s)}}{CuO}$$

b- Detection of Aluminium Al^{+3}:

Salt solution + sodium hydroxide solution gives a white precipitate of Aluminum hydroxide dissolves in excess reagent to form the soluble sodium meta aluminate.

$$\underset{(aq)}{A\ell C\ell_3} + \underset{(aq)}{3NaOH} \longrightarrow \underset{(aq)}{3NaC\ell} + \underset{(s)}{A\ell(OH)_{3\ \text{white PPt}}}$$

$$\underset{(s)}{A\ell(OH)_3} + \underset{(aq)}{NaOH} \longrightarrow \underset{\substack{(aq) \\ \text{sodium metaluminate}}}{NaA\ell O_2} + \underset{(l)}{2H_2O}$$

(2) Sodium Carbonate Na_2CO_3

a- Preparation:

1- **In laboratory** : By passing carbon dioxide gas through a hot solution of sodium hydroxide the solution is left to cool, white crystals of hydrated sodium carbonate are separated gradually.

$$\underset{(aq)}{2Na\,OH} + \underset{(g)}{CO_2} \xrightarrow{\Delta} \underset{(aq)}{Na_2CO_3} + \underset{(l)}{H_2O}$$

The hydrated salt of sodium carbonote is known as washing soda $Na_2CO_3 . 10\,H_2O$. because it's used in removing water hardness which is produced due to presences of Ca^{2+} , Mg^{2+} salts soluble in water, as washing soda reacts with them producting water insoluble calcium and magnesium carbonate so hardness is removed.

$$\underset{(aq)}{Na_2CO_3} + \underset{(aq)}{CaC\ell_2} \longrightarrow \underset{(aq)}{2NaC\ell} + \underset{(s)}{CaCO_3}$$

$$\underset{(aq)}{Na_2CO_3} + \underset{(aq)}{MgSO_4} \longrightarrow \underset{(aq)}{Na_2SO_4} + \underset{(s)}{MgCO_3}$$

2- **In Industry : (Solvay process) :**

By passing ammonia and carbon dioxide gases in a saturated aqueous solution of sodium Chloride to produce sodium bicarbonate . Heating sodium bicarbonate , it will decomposes into sodium carbonate, water and carbon dioxide.

$$\underset{(g)}{NH_3} + \underset{(g)}{CO_2} + \underset{(aq)}{NaC\ell} + \underset{(\ell)}{H_2O} \xrightarrow{heat} \underset{(aq)}{NaHCO_3} + \underset{(aq)}{NH_4C\ell}$$

$$\underset{(aq)}{2NaHCO_3} \xrightarrow{heat} \underset{(aq)}{Na_2CO_3} + \underset{(g)}{CO_2} + \underset{(v)}{H_2O}$$

b- Properties :

1- Whitepowder ,easily dissolves in water . Its solution has an alkaline effect.

2- It is not affected by heat i.e. it melts without decomposition.

3- It reacts with acids , and carbon dioxide evolves .

$$\underset{(s)}{Na_2CO_3} + \underset{(aq)}{2HCl} \longrightarrow \underset{(aq)}{2NaCl} + \underset{(\ell)}{H_2O} + \underset{(g)}{CO_2}$$

c- Uses:

1- Manufacture of glass .

2- Paper industry.

3- Textile industry .

4- Water softening

Biochemical role of sodium:

Sodium ions are from the common ions in blood plasma , and intercellular fluids in the body. It plays an important role in the vital processes, because it forms a suitable medium for transferring nutrients like glucose and amino acids.

From the natural sources of sodium : vegetables (specially celery), milk and its products.

Biochemical role of potassium :

Potassium ions are from the common ions in the cell and plays an important role in the process of oxidation of glucose inside the cell to produce energy needed for it's activity.

From the natural source of potassium : meats , milk , eggs, vegetables and cerials.

Secondly : Elements of p-Block - Group 5 A (Group 15)

This group consists of five elements, they are :

Element	Symbol- At-no	Electronic distribution
Nitrogen	$_7N$	$[He]\ _2\ 2S^2\ 2p^3$
Phosphorus	$_{15}P$	$[Ne]\ _{10}3S^2 3P^3$
Arsenic	$_{33}AS$	$[Ar]\ _{18}4S^2\ 3d^{10}\ 4p^3$
Antimony	$_{51}Sb$	$[Kr]\ _{36}\ 5S^2\ 4d^{10}\ 5p^3$
Bismuth	$_{83}Bi$	$[Xe]\ _{54}\ 6S^2\ 4f^{14}\ 5d^{10}\ 6p^3$

Abundance in nature :

Elements of this group are not abundant, except **nitrogen** which constitutes about 80 % of atmospheric air.($\frac{4}{5}$ of air volume)

is the most abundant element of this group in the earth's crust. It is found in the form of phosphates e.g. calcium Phosphates $Ca_3(PO_4)_2$, apatite CaF_2-$Ca_3(PO_4)_2$.

Arsenic, antimony and bismuth . are found in the form of sulphides as As_2S_3, Sb_2S_3, and Bi_2S_3 .

General properties of Group 5 A elements (Group 15)

The properties of group tend to be those of non metals but metallic properties increase with increasing of atomic number . Nitrogen and phosphorus are non metal, arsenic and antimony are metalloids; but bismuth is a metal, yet its ability to conduct electricity is weak.

2- A nitrogen molecule contains two atoms. But in phosphorus, arsenic and antimony , their vapours have molecules which contain four atoms (P_4,AS_4 ,Sb_4).

-Bismuth forms a metallic crystal lattice, but its vapour consists of diatomic molecules Bi_2. So, it is of an abnormal case ,among all other metals, whose molecules are monoatomic in the vapour state.

3 - Elements of this group are characterized by having several oxidation numbers in their different compound; from (- 3) to (+5). It may gain 3 electrons through covalent sharing, or lose five electrons . The following table shows the oxidation numbers of nitrogen in some of its compounds.

The Compound	Formula	Oxidation no.
ammonia	NH_3	-3
hydrazine	(N_2H_4): NH_2NH_2	-2
hydroxylamine	NH_2OH	-1
nitrogen	N_2	ZERO
nitrous oxide	N_2O	+1
nitric oxide (nitrogen monoxide)	(N_2O_2): NO	+2
nitrogen trioxide	N_2O_3	+3
nitrogen dioxide	(N_2O_4): NO_2	+4
nitrogen pentoxide	N_2O_5	+5

The positive oxidation number are found in oxides, because oxygen has higher electro- negativity than nitrogen.

4 **Allotropy :** "This is the presence of the element in more than one form , of different physical properties, but of similar chemical ones". Solid nonmetals, are characterized by allotropy .The allotropic phenomena is due to the presence of the element in different crystalline forms , each form differs in the number of atoms and in their arrangement.

The element	Allotropic forms
Phosphorus	White (waxy), Red, Violet
Arsenic	Black, grey, yellow (waxy yellow
Antimony	Yellow, Black

Both nitrogen and bismuth , have no allotropic forms .
We shall study in some details nitrogen, since it is the most important element of this group.

5- With oxygen :

all elements of this group from oxides ;Some are acidic N_2O_5, others are amphoteric (Sb_2O_3) or basic.(Bi_2O_5) .The basic property increases with increasing atomic number , and the acidic property decreases with increasing of the atomic number.

6- With hydrogen :

Most of the elements of this group form hydrides (Compounds with hydrogen) , the oxidation number being-3 as in Ammonia (NH_3)Arsine (AsH_3) and PH_3 phosphine . In these compounds, the central atom still has a lone pair of electrons in the valence shell (outer energy level) :It can give this pair of electrons to other atoms or ions making coordinate covalent bonds. Ammonia is more basic than phosphine .The polarity of hydrogen compounds in this group decreases with increasing atomic number . Thus , their solubility in water decreases and their thermal stability decreases so, they decompose even by gentle heating.

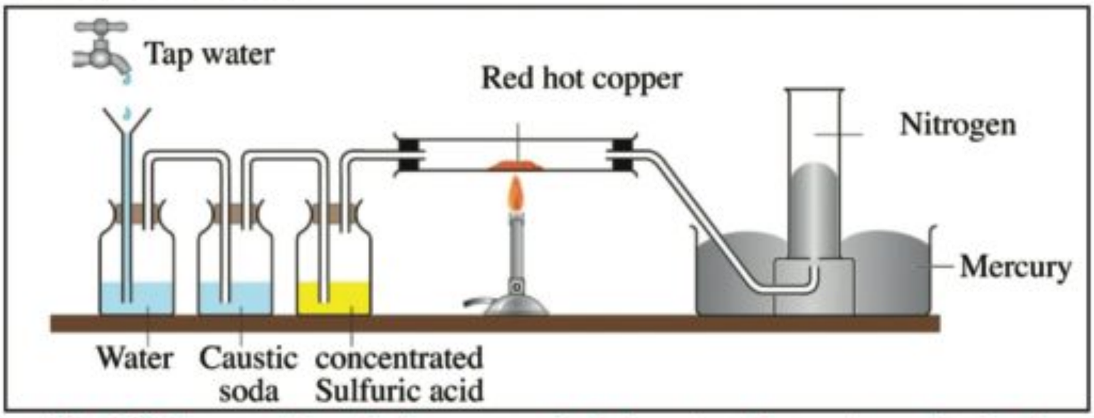
The famous elements in fifth group (A)

Nitrogen (N_2)

Preparation of nitrogen gas in the laboratory :

1) Main method:

It is prepared from atmospheric air (fig 4-3) by removing carbon dioxide, water vapour, and oxygen. and they are the same steps used in preparing the gas in industry commercially.



(fig 4-3) Preparation of nitrogen gas in laboratory from atmospheric air

- As air passes over sodium hydroxide solution to remove (CO_2) gas

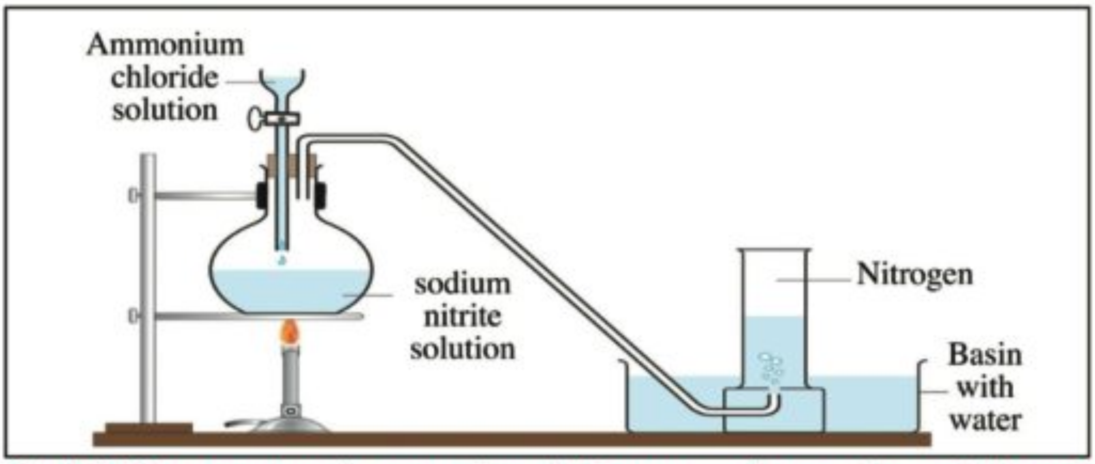
$$[2NaOH_{(aq)} + CO_{2(g)} \longrightarrow Na_2CO_{3(aq)} + H_2O_{(l)}$$

- Then passes over concentration sulpharic acid to absorb water vapour
- then passes over red hot copper turning to remove (O_2) gas

$$2Cu_{(s)} + O_{2(g)} \xrightarrow{\Delta} 2CuO_{(s)}$$

nitrogen gas can be collected by down word displacement of water or on the surface of mercury to get it dry.

2- Nitrogen gas prepared by Heating a mixture of sodium nitrite and ammonium chloride solution



(fig 4-3) The appartus of preparation of nitrogen gas from sodium nitrite and ammonium chloride solution

$$NaNO_{2(aq)} + NH_4C\ell_{(aq)} \xrightarrow{\Delta} NH_4NO_{2(aq)} + NaC\ell_{(aq)}$$

$$NH_4NO_{2(aq)} \xrightarrow{\Delta} 2H_2O_{(\ell)} + N_{2(g)}$$

By adding

$$NaNO_{2(aq)} + NH_4Cl_{(aq)} \xrightarrow{\Delta} NaCl_{(aq)} + 2H_2O_{(\ell)} + N_{2(g)}$$

Physical properties

Practical activity: Physical properties of nitrogen gas:

Prepare several nitrogen gas jars then do the following

Experiment	observation	conclusion
1- What is the colour and odour of the gas? 2- What do you conclude from the way of collecting the gas from atmospheric air, concerning to its solubility in water? 3- Add few drops of pure water to a gas jar containing nitrogen gas and shake them well. Then test the produced solution using (red and blue) litmus paper . What do you notice?		

From the previous activity and others physical properties of nitrogen can be concluded as following :

1- The gas is colourless, odourless and tasteless.
2- The gas is lighter than air because air contain oxygen which is heavier than nitrogen.
3- The gas is sparingly soluble in water (23 ml (N_2) / l (H_2O) at (st.P).
4- Gas has neutral effect on litmus.
5- Its density is (1.25 g/L) at st.p.
6- Its boiling point is (-159 . 79 °C) and can be liquefied at this degree at normal atmospheric pressure.

Properties:

Nitrogen is not an active element .This is due to the strong triple bond which binds the two nitrogen atoms in nitrogen molecule ($N \equiv N$) that is why most of nitrogen reactions takes place in the presence of electric spark, electric arc, or strong heating to break the bond in nitrogen molecule and produce nitrogen atoms.

1- With hydrogen :

by electric spark , ammonia gas is formed (550 °C).

$$\underset{(g)}{N_2} + \underset{(g)}{3H_2} \xrightarrow[\text{electric spark}]{550\ C^\circ} \underset{(g)}{2NH_3}$$

2- With oxygen :

by electric arc (3000 °C) nitric oxide gas is formed . instantly changing into nitrogen dioxide .

$$\underset{(g)}{N_2} + \underset{(g)}{O_2} \xrightarrow[\text{electric arc}]{3000^\circ\ C} \underset{(g)}{2\ NO}$$

Then $\underset{(g)}{2NO} + \underset{(g)}{O_2} \longrightarrow \underset{(g)}{2NO_2}$

3- With metals :

At high temperature nitrogen reacts with metals like magnesium giving the metal nitride .

$$\underset{(s)}{3Mg} + \underset{(g)}{N_2} \xrightarrow{\Delta} \underset{(s)}{Mg_3N_2} \text{ (magnesium nitride)}$$

Nitrides decompose easily in water , ammonia gas evolves .

$$\underset{(s)}{Mg_3N_2} + \underset{(\ell)}{6H_20} \longrightarrow \underset{(g)}{2NH_3} + \underset{(aq)}{3Mg(OH)_2}$$

4- With calcium carbide :

Calcium carbide (CaC_2) reacts with nitrogen, by the electric arc , calcium cyanamide $CaCN_2$ is formed Which is used as a fertilizer.

$$\underset{\text{Calcium Carbide}}{\underset{(s)}{CaC_2}} + \underset{(g)}{N_2} \xrightarrow[3000^\circ\ C]{\text{electric arc}} \underset{\text{Calcium cyanamide}}{\underset{(s)}{CaCN_2}} + \underset{(s)}{C}$$

Calcium cyanamide is a source of ammonia in the agricultural soil while land is being irrigated.

$$\underset{(s)}{CaCN_2} + \underset{(\ell)}{3H_2O} \longrightarrow \underset{(s)}{CaCO_3} + \underset{(g)}{2NH_3}$$

The most famous nitrogenous compounds

(Firstly) Ammonia NH_3

Practical Experiment: Preparation of ammonia in the lab :

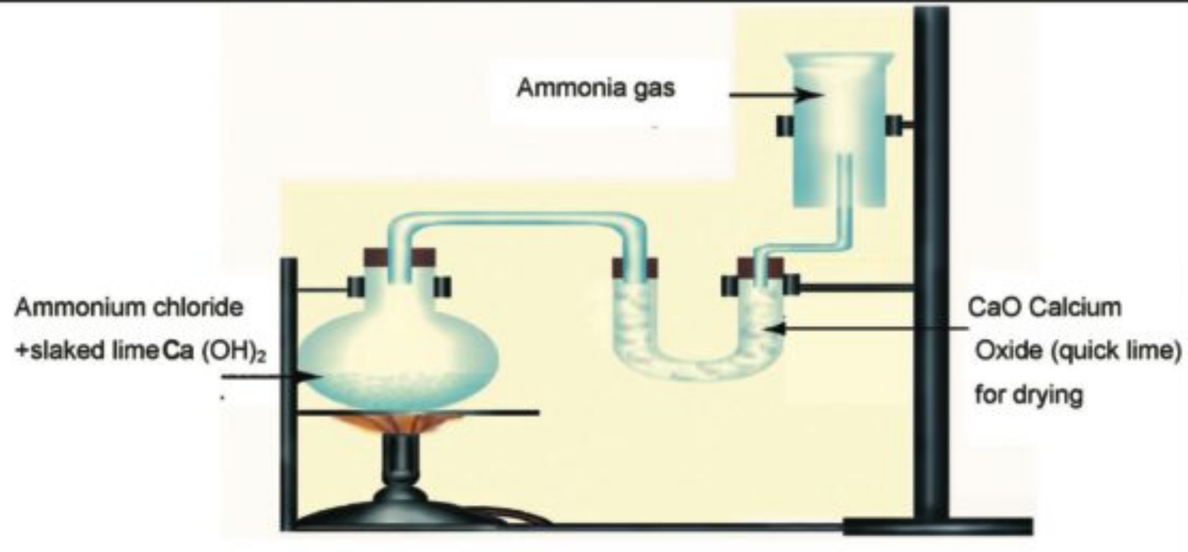
1- Set the apparatus as in figure (4-5)

2- Put in the flask a mixture of ammonium chloride and slaked lime .

3- Put a drying agent (quick lime CaO) in a U- shaped tube.

4- Heat the contents of the flask, then fill cylinders by down word displacement of air. Test the properties of the gas collected.

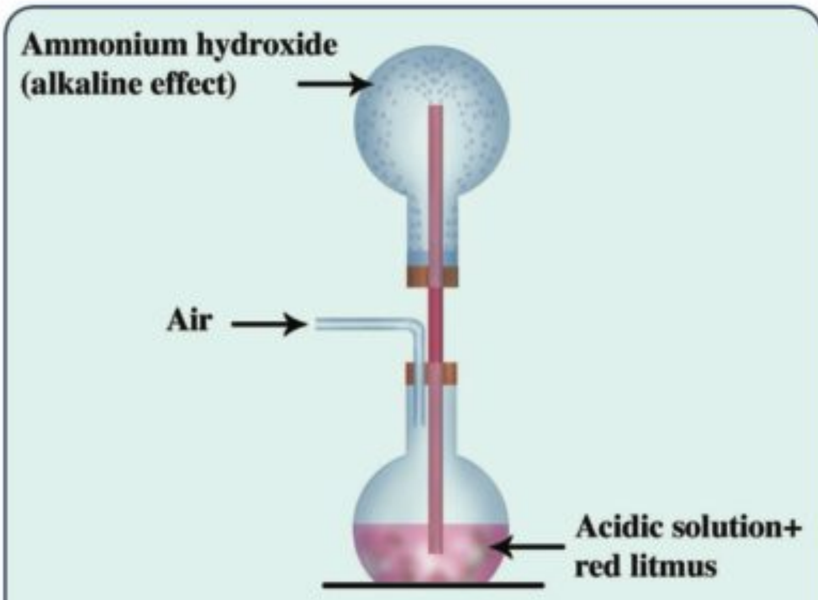
$$\underset{(s)}{2NH_4Cl} + \underset{(s)}{Ca(OH)_2} \xrightarrow{\Delta} \underset{(s)}{CaCl_2} + \underset{(\ell)}{2H_2O} + \underset{(g)}{2NH_3}$$



Preparation of ammonia in lab

(fig 4- 5)

Experiment	Observation	Conclusion
1- What is the colour and odour of ammonia?		
2- Put few drops of water in the cylinder full of ammonia and shake. Then add red litmus solution to the cylinder. What is the action of the gas on it?		
3- Introduce a glowing splint near the top of the cylinder . Is the glowing splint extinguished or does the gas burn?		
4- Set the apparatus of the fountain experiment as in the figure (4- 6). fill the bottle with ammonia gas. Notice what happens. Why?		



(fig 4- 6)
Fountain exp.

Why is ammonia considered as a base anhydride?

Preparation of ammonia gas in industry

Haber Bosh method :

Ammonia gas can be prepared in industry from nitrogen and hydrogen, in the presence of a catalyst: Iron, molybdenum , and under 200 atmospheric pressure and 500°C .

$$\underset{(g)}{N_2} + \underset{(g)}{3\,H_2} \xrightarrow[\text{500 C 200 atm.p}]{\text{fe/Mo catalyst}} \underset{(g)}{2\,NH_3}$$

Detection of ammonia gas:

Ammonia gas forms a dense white fumes of ammonium chloride (solid substance sublimes) when subjected to a glass rod wetted with conc. hydrochloric acid .

$$\underset{(g)}{NH_3} + \underset{(g)}{HC\ell} \longrightarrow \underset{(s)}{NH_4C\ell}$$

The role of ammonia in the manufacture of fertilizers

Ammonia is synthesized in industry from nitrogen .Nitrogen is considered one of the main sources of nutrition in plants since it forms protein . Nitrogen is present in the soil in the form of inorganic and organic substances , the amount of nitrogen in the soil decreases by time.

The soil should be enriched by nitrogen either by adding synthetic nitrogenous fertilizers or natural fertilizers (manure) . Although nitrogen forms about 4/5 of the air volume plant can not use it by a direct way in its gaseous form . Accordingly ,it is a must to provide the soil by nitrogen in the form of ammonium salts or urea which dissolve in the irrigation water where it is absorbed by the plant roots.

Ammonia is considered the starting material for most of the nitrogenous fertilizers.

Synthesis of some important Nitrogenous fertilizers

1- Ammonium salts are usually used as inorganic fertilizers , They are synthesized by a neutralization reaction between ammonia and the suitable acid .

$$\underset{(g)}{NH_3} + \underset{(\ell)}{HNO_3} \longrightarrow \underset{(aq)}{NH_4NO_3} \text{ ammonium nitrate}$$

$$\underset{(g)}{2NH_3} + \underset{(aq)}{H_2SO_4} \longrightarrow \underset{(aq)}{(NH_4)_2SO_4} \text{ Ammonium Sulphate}$$

2- To prepare a nitrogenous phosphorous fertilizer phosphoric acid is prepared first, then it reacts with ammonia.

$$\underset{(aq)}{H_3PO_4} + \underset{(g)}{3NH_3} \longrightarrow \underset{(aq)}{(NH_4)_3PO_4} \text{ Ammonium, phosphate}$$

Notes on fertilizers :

1- Ammonium nitrate has a high nitrogen content (35 %).It is highly soluble in water. Excess amount of ammonium nitrate causes acidic effect on soil.

2- The continuous use of ammonium sulphate as a fertilizer increases acidity of soil. Thus it is important to neutralize the soil by basic substance.

3- Ammonium phosphate is an important fertilizer because it provides the soil with two essential elements which are nitrogen and phosphorous.

4- Urea has a high nitrogen content (46 %).It is the most suitable fertilizer for hot areas , since the high temperature increases its dissociation to ammonia and carbon dioxide .

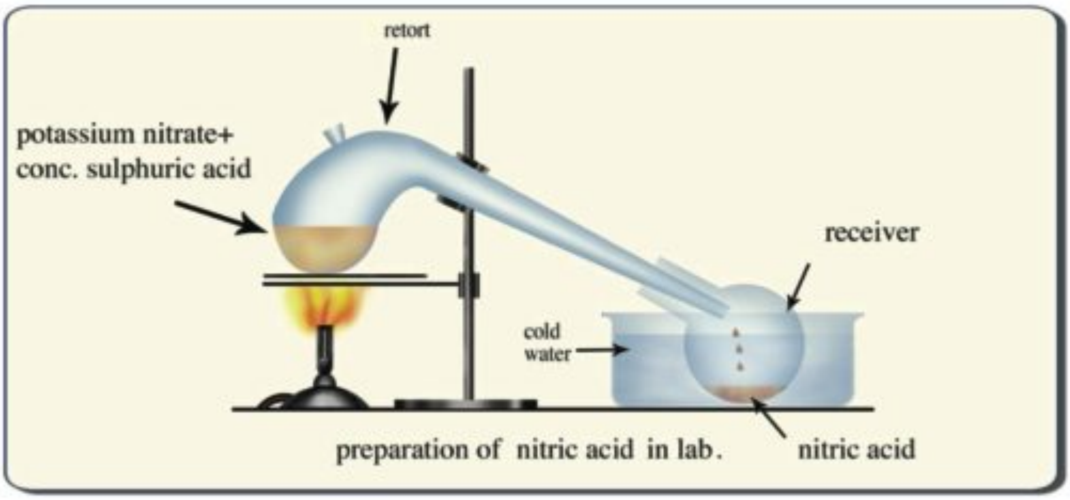
5- Anhydrous liquefied ammonia is the future fertilizer because it has the highest nitrogen content over all fertilizers (82 %).It is supplied to the soil in a depth of nearly 12 cm .

Secondly :Nitric Acid (HNO_3)

Preparation of nitric acid in the lab.

Practical experiment:

1- Set the apparatus as in figure (4 - 7).



(fig 4- 7)

2- Put potassium nitrate and concentrated sulphuric acid in the retort and the receiver in a trough containing cold water.

3- Heat the contents of the retort gently , taking care that the temperature does not exceed 100° C.

Collect the acid formed in the receiver. Test the acid.

$$\underset{(s)}{2KNO_3} + \underset{(aq)}{H_2SO_4} \xrightarrow[\text{Heat}]{\text{Conc}} \underset{(aq)}{K_2SO_4} + \underset{(\ell)}{2HNO_3}$$

Experiment	Observation	Conclusion
1. What is the colour of the liquid formed? 2- Add blue litmus solution to the liquid formed. What do you see ? 3- In a test tube, place a small volume of the acid solution. Add to it Cu turning then heat. What do you observe? 4- Dilute some of the acid with water, then add iron filling . What is the colour of the gas evolved? What is the action of air on it?		

Chemical properties of nitric acid :

1- **Action of heat:** decomposes by heat giving brown fumes of nitrogen dioxide

$$\underset{(\ell)}{4HNO_3} \xrightarrow{\Delta} \underset{(g)}{4NO_2} + \underset{(g)}{O_2} + \underset{(\ell)}{2H_2O}$$

2- **It is oxidizing agent:** because oxygen gas evolved as a result of is thermal decomposition .

3- Action of the acid on metals

The acid reacts with metals above hydrogen in the electrochemical series , giving metal nitrate and atomic hydrogen, which reduces the acid.

Example of reaction between some metals and the acid:

$$\underset{(s)}{Fe} + \underset{(aq)}{4HNO_3} \xrightarrow[\Delta]{dil} \underset{(aq)}{Fe(NO_3)_3} + \underset{(l)}{2H_2O} + \underset{(g)}{NO}$$ Nitric Oxide

Metals below hydrogen in the electrochemical series react with the acid which acts as an oxidizing agent; the metal is oxidized then the oxide reacts with the acid.

$$\underset{(s)}{3\,Cu} + \underset{(aq)}{8HNO_3} \xrightarrow[\Delta]{dil} \underset{(aq)}{3Cu(NO_3)_2} + \underset{(g)}{2NO} + \underset{(\ell)}{4H_2O}$$

$$Cu_{(s)} + 4HNO_{3(l)} \xrightarrow{Conc} Cu(NO_3)_{2(aq)} + 2NO_{2(g)} + 2H_2O_{(\ell)}$$

1. Some metals e.g. Fe, Cr, $A\ell$ are not affected by the concentrated acid. This is due to **passivity** .

Passivity of the metal is due to the oxidizing property of the acid; a layer of the metal oxide is formed, this layer is non porous so it protects the metal from further reaction.

Detection of nitrate ion NO_3^-

The brown ring experiment:

1- To a nitrate salt solution add a freshly prepared conc. solution of iron (II) sulphate.

2- Add carefully on the internal walls of the tube few drops of conc.H_2SO_4.

A Brown ring appears at interface and disappears by shaking or heating.

$$2NaNO_{3(aq)} + 6FeSO_{4(aq)} + 4H_2SO_{4(aq)} \longrightarrow 3Fe_2(SO_4)_{3(aq)} + Na_2SO_{4(aq)} + 4H_2O_{(l)} + 2NO_{(g)}$$

$$FeSO_{4(aq)} + NO_{(g)} \longrightarrow FeSO_4.NO_{(s)}$$ brown ring compound

How to differentiate between nitrite and nitrate salts?

Add to the salt solution a solution of potassium permanganate acidified with conc. sulphuric acid . The colour disappear in case of nitrite salt.

$$5KNO_{2(aq)} + 2KMnO_{4(aq)} + 3H_2SO_{4(aq)} \xrightarrow{Conc} 5KNO_{3(aq)} + K_2SO_{4(aq)} + 2\,MnSO_{4(aq)} + 3\,H_2O_{(\ell)}$$

Nitrate salt does not remove the colour of permanganate .

Give reason: it is impossible to perform the flame test for elements of group 5A except (Sb , As)

Economic importance of group 5A elements

1- Nitrogen:

- Ammonia gas industry.
- Nitrogenous fertilizer industry.
- Filling car tyres (because it decreases the probability of exploding tyres, due to its low response to the temperature change in addition to its lower rate of leakage than atmospheric air.
- Filling chips bags to keep potato's slices crunchy due to its relative inactivity.
- Liquefied nitrogen used in preservation and transferring the living cells and treatment of banging tumers.

2- Phosphorus:

- To make matches.
- Phosphorus fertilizers industry.
- To make alloys like phosphorus bronze alloy (Cu - Sn - P) to make pushing fans for ships .
- To make fire works.

3- Arsenic: (highly toxic):

- Used as a preservative substance for woods due its poisonious effect on insects , bacteria and fungi.
- To make arsenic trioxide used in the treatment of (leukemia)

4- Antemoni:

- To make lead antemonal alloy used in car batteries.
- Used in making semiconductors that used in making instruments of detecting I.R.

5- Bismoth:

- To make alloy with lead and cadmium used in electric fuses due its low melting point.

QUESTION 1:-

Explain the electronic configuration by building up principle of each of the following elements ,then show their possible oxidation states :

Potassium (K_{19}) - Caesium (CS_{55}) - Nitrogen (N_7) Phosphorus (P_{15}) .

QUESTION 2:-

Give reasons for :

1- The chemical reactivity of alkali metals.

2- The weakness of metallic bonds between the atoms of the first group elements.

3- The use of Caesium in photo-electric cells.

4- The use of potassium super oxide in submarines.

5- Water is not used to extinguish sodium fires .

6- Sodium nitrate is not used as an explosive substance.

7- It is difficult to extract the alkali metals from their ores by the normal chemical methods.

8- Nitrogen has several oxidation states .

9- Calcium cyanamide is used as a fertilizer.

10- The alloy of bismuth, lead, cadmium and tin is used as fuses.

11- Nitric acid is considered a strong oxidizing agent.

12- Urea is used as a fertilizer in hot countries .

13- Conc. hydrochloric acid is used in the detection of ammonia gas.

14- Its more preferable to fill car tyres with nitrogen gas instad of atmospheric air.

15- Arsenic used as preservative substance for woods.

QUESTION 3
How to differentiate practically between:

a- Sodium nitrate and sodium nitrite .

b- Copper sulphate and aluminium sulphate

QUESTION 4:-
Explain how sodium carbonate is prepared in industry.

QUESTION 5:-
Explain the effect of each of the following on sodium metal.

Hydrochloric acid - Hydrogen - oxygen -water

QUESTION 6:-
Explain how nitric acid is prepared in laboratory. Write the balanced equation and draw the used apparatus .

QUESTION 7:-
Choose the correct answer.

a- The metallic property in the first group elements increases by increasing ....

1- the abundance in the earth crust. 2- the boiling point.

3 - Atomic number . 4- melting point.

b- In the vapour state phosphorus molecule contains...............

1- one atom 2- two atoms

3 - three atoms 4-four atoms

c- When calcium cyanamide reacts with water , the evolve gas is ....

1- ammonia 2-Hydrogen

3 - nitric oxide 4- nitrogen dioxide

d- The oxidation no .of the first group elements in their compounds is ....

1-(-1) 2- (+ 1) 3-(-2) 4-(+ 2)

e- Potassium Super oxide is used in submarines to exchange carbon dioxide by ....

1- Hydrogen
2 - Oxygen
3 - Ammonia
4 - Carbon monoxide

f - When a glass rod wetted with conc. hydrochloric acid is exposed to ammonia gas , a dense white clouds of .... is formed .

1- Ammonium carbonate
2- Ammonium chloride
3 - Hydrogen chloride
4-Ammonium sulphate,

g - On adding sodium hydroxide to copper sulphate solution then heating the product, the colour of the precipitated substance is ...

1-black
2-white
3-yellow
4-red

QUESTION 8:-

Explain with drawing a method of preparation of nitrogen gas from atmospheric air. Write labels on the drawing with writing chemical equation takes place.

جميع حقوق الطبع محفوظة لوزارة التربية والتعليم
داخل جمهورية مصر العربية

المواصفات الفنيــة:

مقـــاس الكتـــــاب:	$\frac{1}{8}$ (٥٧ × ٨٢) سم
طبــــع الـمتــــن:	٤ لون
طبـــــع الغــلاف:	٤ لون
ورق الـمتـــــــن:	٨٠ جم أبيض
ورق الغـــــــلاف:	٢٠٠ جم كوشيه
عدد الصفحات بالغلاف:	١٢٠ صفحة

الشروقـــــ
الحديثة للطباعه والتغليف